FBISE

CHEMISTRY

MODEL PAPERS & GUESS PAPERS

Faderal Board Islamabad

9TH CLASS

Presented by:

Urdu Books Whatsapp Group

STUDY GROUP

0333-8033313	0343-7008883	0306-7163117
راؤ ایاز	پاکستان زندہ باد	محمد سلمان سلیم

GUESS PAPER & MODEL PAPER # 1
BASED ON CHAPTER # 1 (Reduced Syllabus)
FUNDAMENTALS OF CHEMISTRY

CHAPTER 1: BRANCHES OF CHEMISTRY

Basic definitions, Matter, substance, Mixture(types), Elements, compound, Atomic number, Mass number, Relative mass unit, Empirical formula and Molecular formula, Types of molecular formula, Molecular mass calculation, concept of mole, Avogadro's number, chemical calculation, mole-mass calculation (1.5), Mole-particles calculation Example (1.6,1.7)

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

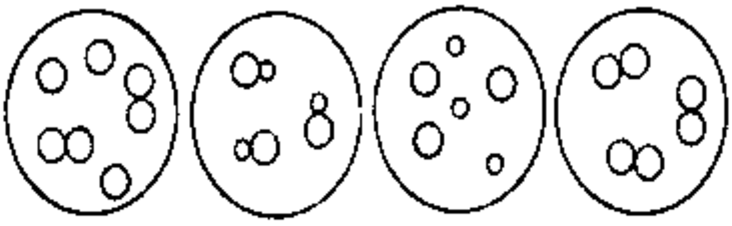
Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) How many numbers of moles are equivalent to 8 grams of CO_2?
A. 0.15 B. 0.18 C. 0.21 D. 0.24

(ii) The diagrams below represent particles in four substances, which box represent the particles in nitrogen.



A. B. C. D.

(iii) What is the formula mass of $CuSO_4.5H_2O$. (Atomic masses: Cu=63.5, S=32, O=16, H=1)
A. 159.5 B. 185.5 C. 249 5 D. 149.5

(iv) A compound with chemical formula Na_2CX_3 has formula mass 106amu. Atomic mass of the element X is.
A. 106 B. 23 C. 12 D. 16

(v) How many moles of molecules are there in 16 g oxygen?
A 1 B. 0.5 C 0 1 D. 0.05

(vi) What is the mass of 4 moles of hydrogen gas?
A. 8.064 g B. 4.032 g C. 1 g D. 1.008 g

(vii) What is the mass of carbon present in 44 g of carbon dioxide?
A. 12 g B. 6 g C. 24 g D. 44 g

(viii) The electron configuration of an element is $1s^2 2s^2$. An atom of this element will form an ion that will have charge.

(x) If one mole of carbon contains x atoms, what is the number of atoms contained in 12 g of Mg.
A. x B. $0.5x$ C. $2x$ D. $1.5x$

(xi) Relationship between the Empirical and Molecular Formula:
A. Molecular formula = 2n × Empirical formula
B. Molecular formula = n^2 × Empirical formula
C. Molecular formula = n × Empirical formula
D. Molecular formula = $2n^2$ × Empirical formula

(xii) Which of the following lists contains only elements?
A. Air, water, oxygen B. Hydrogen, oxygen, brass
C. Air, water, fire, earth D. Calcium, sulphur, carbon

CHEMISTRY SSC-I

Time allowed: 2:40 hours **Total Marks: 53**

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet–B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. What is mole?
ii. Differentiate between empirical formula and molecular formula.
iii. What is the number of molecules in 9.0 g of steam?
iv. What are the molar masses of uranium -238 and uranium -235?
v. Why one mole of hydrogen molecules and one mole of H-atoms have different masses?
vi. Differentiate between (a) atom and ion (b) molecular ion and free radical.
vii. Describe how Avogadro's number is related to a mole of any substance.
viii. Decide whether or not each of the following is an example of empirical formula:
a. Al_2Cl_6 b. Hg_2Cl_2 c. NaCl d. C_2H_6O

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. Define compound.
ii. Define mixture.
iii. Describe the different types of mixture?
iv. Differentiate between compound and mixture?
v. Define molecular mass.
vi. Define formula mass.
vii. Define molar mass.

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q4. a. Define ion, molecular ion, formula unit, free radical, atomic number, mass number, atomic mass unit.
b. Calculate the total number of atoms present in 18g H_2O.

Q5. a. Define matter, mass, substance, element, atomic number and mass number.
b. When natural gas burns, CO_2 is formed. If 0.25 moles of CO_2 is formed, what mass of CO_2 is produced?

Q6. a. Define Avogadro's number. How does it relate to the masses of chemical

ختم نبوت ﷺ زندہ باد

عظمت صحابہ زندہ باد

السلام علیکم ورحمۃ اللہ وبرکاتہ:

معزز ممبران: آپ کا وٹس ایپ گروپ ایڈمن "**اردو بکس**" آپ سے مخاطب ہے۔

آپ تمام ممبران سے گزارش ہے کہ:

- گروپ میں صرف PDF کتب پوسٹ کی جاتی ہیں لہذا کتب کے متعلق اپنے کمنٹس / ریویوز ضرور دیں۔ گروپ میں بغیر ایڈمن کی اجازت کے کسی بھی قسم کی (اسلامی وغیر اسلامی، اخلاقی، تحریری) پوسٹ کرنا سختی سے منع ہے۔
- گروپ میں معزز، پڑھے لکھے، سلجھے ہوئے ممبرز موجود ہیں اخلاقیات کی پابندی کریں اور گروپ رولز کو فالو کریں بصورت دیگر معزز ممبرز کی بہتری کی خاطر ریموو کر دیا جائے گا۔
- کوئی بھی ممبر کسی بھی ممبر کو انباکس میں میسیج، مس کال، کال نہیں کرے گا۔ رپورٹ پر فوری ریموو کر کے کاروائی عمل میں لائے جائے گی۔
- ہمارے کسی بھی گروپ میں سیاسی و فرقہ واریت کی بحث کی قطعاً کوئی گنجائش نہیں ہے۔
- اگر کسی کو بھی گروپ کے متعلق کسی قسم کی شکایت یا تجویز کی صورت میں ایڈمن سے رابطہ کیجئے۔
- سب سے اہم بات:

گروپ میں کسی بھی قادیانی، مرزائی، احمدی، گستاخِ رسول، گستاخِ امہات المؤمنین، گستاخِ صحابہ و خلفائے راشدین حضرت ابو بکر صدیق، حضرت عمر فاروق، حضرت عثمان غنی، حضرت علی المرتضیٰ، حضرت حسنین کریمین رضوان اللہ تعالیٰ اجمعین، گستاخ اہلبیت یا ایسے غیر مسلم جو اسلام اور پاکستان کے خلاف پراپیگنڈا میں مصروف ہیں یا ان کے روحانی و ذہنی سپورٹرز کے لئے کوئی گنجائش نہیں ہے لہذا ایسے اشخاص بالکل بھی گروپ جوائن کرنے کی زحمت نہ کریں۔ معلوم ہونے پر فوراً ریموو کر دیا جائے گا۔

- تمام کتب انٹرنیٹ سے تلاش / ڈاؤنلوڈ کر کے **فری آف کاسٹ** وٹس ایپ گروپ میں شیئر کی جاتی ہیں۔ جو کتاب نہیں ملتی اس کے لئے معذرت کر لی جاتی ہے۔ جس میں محنت بھی صَرف ہوتی ہے لیکن ہمیں آپ سے صرف دعاؤں کی درخواست ہے۔
- عمران سیریز کے شوقین کیلئے علیحدہ سے عمران سیریز گروپ موجود ہے۔
- **لیڈیز کے لئے الگ گروپ کی سہولت موجود ہے جس کے لئے ویریفکیشن ضروری ہے۔**
- اردو کتب / عمران سیریز یا سٹڈی گروپ میں ایڈ ہونے کے لئے ایڈمن سے وٹس ایپ پر بذریعہ میسیج رابطہ کریں اور جواب کا انتظار فرمائیں۔ برائے مہربانی اخلاقیات کا خیال رکھتے ہوئے موبائل پر کال یا ایم ایس کرنے کی کوشش ہر گز نہ کریں۔ ورنہ گروپس سے تو ریموو کیا ہی جائے گا بلاک بھی کیا جائے گا۔

نوٹ: ہمارے کسی گروپ کی کوئی فیس نہیں ہے۔ سب فی سبیل اللہ ہے

0306-7163117	0343-7008883	0333-8033313
محمد سلمان سلیم	پاکستان زندہ باد	راؤ ایاز

پاکستان زندہ باد

پاکستان پائندہ باد

اللہ تبارک تعالیٰ ہم سب کا حامی وناصر ہو

(a) A balloon filled with 5g of hydrogen.
(b) A block of ice that weighs 100g.

SOLUTION OF GUESS PAPER & MODEL PAPER # 1
(Reduced Syllabus)

SECTION- A (MCQs)

i. B	ii. A	iii. C	iv. D	v. B	vi. A
vii. A	viii. B	ix. D	x. B	xi. C	xii. D

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. What is mole?

Ans: Mole Concept:

The amount of matter that contains as many atoms ions or molecules as the number of atoms in exactly 12g of C-12 is called mole.

Mole can also be defined as atomic mass, molecular mass or formula mass expressed in grams

$$\text{Number of moles of a substance} = \frac{\text{Mass in grams}}{\text{Molecular mass}}$$

A mole is an amount of a substance that contains 6.02×10^{23} particles of that substance. This experimentally determined number is known as Avogadro's number

Definition of mole:

The definition of mole, accepted internationally, is as follows:

Mole is the amount of a substance which contains the same number of chemical units as there are atoms in exactly 12 grams of pure carbon-12.

Symbol of the mole unit:

The unit of mole is given a symbol 'mol'.

Example:

For example a mole of carbon is 6.02×10^{23} atoms A mole of sulphur is 6.02×10^{23} atoms. A mole of water is 6.02×10^{23} molecules.

ii. Differentiate between empirical formula and molecular formula.

Ans: Difference between empirical and molecular formula:

Empirical Formula	Molecular formula
1. The empirical formula of a compound is the chemical formula that gives the simplest whole-number ratio of atoms of each element.	1. A molecular formula gives the actual whole number ratio of atoms of each element present in a compound.
2. A formula which represents the simplest whole number ratio of atoms of elements in a compound.	2. A formula which represents actual number of atoms of elements in a compound.
3. It is obtained from elemental analysis. $Empirical\ Formula = \frac{Molecular\ Formula}{n}$	3. It is obtained by multiplying 'n' with empirical formula where: $n = \frac{Molecular\ Mass}{Empirical\ Formula\ Mass}$
4. It is used for both molecular and ionic compounds.	4. It is used for molecular compounds.
5. Examples: CH_2O and CH are empirical formulae of glucose and benzene respectively.	**5. Examples:** $C_6H_{12}O_6$ and C_6H_6 are molecular formulae of glucose and benzene respectively.

iii. What is the number of molecules in 9.0 g of steam?

Avogadro's number = $N_A = 6.022 \times 10^{23}$
Number of Molecules = ?

Number of molecules = $\frac{\text{Mass in grams}}{\text{Molar mass}} \times N_A$

Number of molecules = $\frac{9}{18} \times 6.022 \times 10^{23}$

Number of molecules = $0.5 \times 6.022 \times 10^{23}$

Number of molecules = 3.011×10^{23} Molecules

iv. What are the molar masses of uranium -238 and uranium -235?

Ans: Molar mass of uranium -238 = Atomic mass of Uranium 238 = 238 g
Molar mass of uranium -235 = Atomic mass of Uranium -235 = 235 g

v. Why one mole of hydrogen molecules and one mole of H-atoms have different masses?

Ans: One-mole quantities of two different substances have different masses for the same reason—the substances have different compositions.

If we put one mole of hydrogen molecules and one mole of H-atoms on separate balances, we would see a difference in mass, just as you do for the eggs and the limes. This occurs because hydrogen molecules differ from H-atoms. Thus, the mass of 6.02×10^{23} hydrogen molecules does not equal the mass of 6.02×10^{23} H-atoms. (1 mole of H-atom = 1g , 1 mole of H_2 = 1 × 2 = 2 g)

vi. Differentiate between (a) atom and ion (b) molecular ion and free radical.

Ans: (a) Difference between atom and ion:

An ion makes up the electric charge of an atom. It can be a positively (+) charged atom or a negatively (-) charged atom, depending on the number of protons versus electrons.

On the other hand, an atom is the smallest part of an element composed of electrons, protons, and the nucleus.

(b) Difference between molecular ion and free radical:

Polyatomic and molecular ions are often formed by the combination of elemental ions such as H^+ with neutral molecules or by the loss of such elemental ions from neutral molecules. Many of these processes are acid-base reactions.

A radical ion is a free radical species that carries a charge. Radical ions are encountered in organic chemistry as reactive intermediates.

vii. Describe how Avogadro's number is related to a mole of any substance.

Ans: Relation of Avogadro's number to mole:

One mole of a substance = 6.02×10^{23} atoms / molecules

$$\text{Number of moles of a substance} = \frac{\text{Number of molecules of substance}}{N_A}$$

$$\text{Number of moles of a substance} = \frac{\text{Number of molecules of substance}}{6.02 \times 10^{23}}$$

Examples: One mole of hydrogen atoms = 6.02×10^{23} atoms of hydrogen (H) = 1 a.m.u
One mole of hydrogen molecules = 6.02×10^{23} molecules of hydrogen (H_2) = 2 a.m.u
One mole of water = 6.02×10^{23} molecules of water (H_2O) = 18 a.m.u

viii. Decide whether or not each of the following is an example of empirical formula:

a. Al_2Cl_6 **b.** Hg_2Cl_2 **c.** NaCl **d.** C_2H_6O

Solution: a. Al_2Cl_6

No, since, 2 : 6 is not the simplest whole number ratio therefore Al_2Cl_6 is not empirical formula. Hence, Al_2Cl_6 is a molecular formula.

b. Hg_2Cl_2:

No, since, 2 : 2 is not the simplest whole number ratio therefore Hg_2Cl_2 is not empirical formula. Hence, Hg_2Cl_2 is a molecular formula.

c. **NaCl:** Yes, since, 1 : 1 is the simplest whole number ratio therefore NaCl is empirical formula.

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. **(5 × 3 = 15)**

i. Define compound.

Ans: compound:

A compound consists of two or more elements held together in fixed proportions by chemical bonds.

The properties of compounds are different from the properties of the elements from which they are formed.

Examples: Water, carbon dioxide, copper sulphate, sodium chloride etc are compounds. Elements and compounds have uniform composition throughout

ii. Define mixture.

Ans: Mixture:

An impure substance that contains two or more pure substances that retain their individual chemical characteristics is called a mixture. A mixture can be converted into two or more pure substances by physical method.

Examples: Examples of mixture are air, water containing dissolved oxygen, table salt dissolved in water, salt + sand etc

iii. Describe the different types of mixture?

Ans: Types of mixture:

Heterogeneous mixture:

A mixture that consists of two or more visibly different components is called a heterogeneous mixture.

Examples: Sand + salt, oil floating on water etc.

Homogeneous mixture: A mixture consists of only one phase is called a homogeneous mixture.

Examples: Sugar mixed in water, salt dissolved in water Homogeneous mixtures also have uniform composition throughout.

Note: In fact, the entire physical world is made up of mixture of elements and compounds. Most of its components are made up of molecules.

iv. Differentiate between compound and mixture?

Ans: Distinction between compound and mixture:

Following are the points of difference between a compound and a mixture:

No.	Compound	Mixture
1.	A compound consists of two or more elements held together in fixed proportions by chemical bonds.	An impure substance that contains two or more pure substances that retain their individual chemical characteristics is called a mixture.
2.	All compounds have fixed melting and boiling points.	Melting and boiling points of any element or compound present in the mixture are quite different to the real ones.
3.	The elements forming the compounds cannot be separated by either physical or mechanical means but with chemical means only.	The elements or the compounds making the mixture can be easily separated by physical or mechanical means.
4.	When gases react together at constant temperature and pressure to form a compound change in volume usually takes place	When gases are mixed together at constant temperature and pressure to form a gaseous mixture, no change in volume takes place.
5.	The properties of the elements making the compound are quite different to those of the compound.	The properties of the mixture are between the properties of the constituents forming the mixture.
6.	Hydrogen chloride (HCl), glucose ($C_6H_{12}O_6$), carbon dioxide (CO_2), etc are compounds	(i) Air is a mixture of nitrogen (N_2), oxygen (O_2), carbon dioxide (CO_2) etc. (ii) Gun powder is a mixture of nitre, sulphur

v. Define molecular mass.

Ans: Molecular mass:

Molecular mass is the sum of atomic masses of all the atoms present in the molecule

Examples: (i) **Molecular mass of water (H_2O):**

Molecular mass of water H_2O = 2(atomic mass of H) + atomic mass of oxygen
= 2(1.008) + 16.00 = 2.016 + 16.00 = 18.016 amu

(ii) **Molecular mass of sulphuric acid (H_2SO_4):**

Molecular mass of H_2SO_4 = (2) + (1 × 32) + (4 × 16) = 98 amu

vi. Define formula mass.

Ans: Formula mass:

The sum of the atomic masses of all the atoms in the formula unit of a substance is called formula mass.

Examples:

1. Formula mass of NaCl = 1 × Atomic mass of Na + 1 × Atomic mass of Cl = 1 × 23 + 1 × 35.5
= 58.5 amu

2. Formula mass of $Mg(OH)_2$ = 24 + 16×2 + 1 × 2 = 24 + 32 + 2 = 58 amu

Note: The term molecular mass is used for molecular compounds. Whereas, the term formula mass is used for ionic compounds.

vii. Define molar mass.

Ans: Molar mass: The mass of one mole of substance is called as molar mass.

Examples: For instance water exists as molecules, therefore, one mole of water contains 6.02×10^{23} molecules of water. Hydrogen exists as H_2 molecules, so one mole of hydrogen contain 6.02×10^{23} molecules. Carbon exists as atoms so 1 mole of carbon contains 6.02×10^{23} atoms

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. **(2 × 10 = 20)**

Q4. a. Define ion, molecular ion, formula unit, free radical, atomic number, mass number, atomic mass unit.

Ans: Ion: Ion is a charged species formed from an atom or chemically bonded groups of atoms by adding or removing electrons.

Cations: Positively charged ions are called cations.

Anions: Negatively charged ions are called anions.

Examples: For example Na forms Na^+ by losing one electron, **Ca** forms Ca^{+2} by losing two electrons. The Non-metal atoms usually gain one or more electrons and form anions.

Examples: For example chlorine atom gains one electron and forms Cl^- ion, O-atom gains two electrons and forms O^{-2} ion.

Molecular ion: When a molecule loses or gains electrons, the resulting species is called a molecular ion.

Example: For example O_2 when loses one electron it forms O_2^+ ion, but when it absorbs an electrons it forms O_2^- ion. These ions are called molecular ions.
Similarly, N_2^-, N_2^+ etc are examples of molecular ions.

Note: These are short lived species and only exist at high temperature. Molecular ions do not form ionic compounds.

Formula unit: Formula unit is the simplest formula of an ionic compounds
A formula unit, as the name implies, is one unit, where atom, ion or molecule corresponds to given formula.

For example: One formula unit of NaCl has one Na^+ ion and one Cl^- ion. One formula unit of $MgBr_2$ has one Mg^{2+} ion and two Br^- ions.

Free radicals: A free radical is an atom, which has an unpaired electron and bears no electrical charge

For Example: $\dot{H}$, $:\ddot{Cl}\cdot$ are free radicals

Note: When substances like halogens are exposed to sun light, their molecules split up into free radicals.

Mass number: The total number of protons and neutrons in an atom is known as its mass number.
Number of neutrons = mass number – atomic number

Atomic mass unit: One atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one C-12 atom. **Mass of one C-12 atom = 12 amu.**

$$1\,\text{amu} = \frac{\text{mass of one C}-12\text{ atom}}{12}$$

A hydrogen atom is 8.40% as massive as the standard C-12 atom. Therefore, relative atomic mass of hydrogen

$$= \frac{8.40}{100} \times 12\,\text{amu} = 1.008\,\text{amu}$$

Similarly, relative atomic masses of O, Na, Al are 15.9994 amu, 22.9898 amu, 26.9815 amu respectively.

b. Calculate the total number of atoms present in 18g H_2O.

Ans: Solution:

Known mass of H_2O = 18 g = 1 mole of H_2O = 6.02×10^{23} atoms.
Since one molecule of water has two atoms of hydrogen, and one atom of oxygen.
Therefore, one molecule of water has total number of atoms = 2 + 1 = 3 atoms.
Total number of atoms in 18 g of water = $3 \times 6.02 \times 10^{23} = 1.806 \times 10^{24}$ atoms.

Q5. a. Define matter, mass, substance, element, atomic number and mass number.

Ans: Matter: Anything that occupies space and has mass is called matter.

Mass: Quantity of matter in a body is called its mass.

Substance: Any matter that has a particular set of characteristics that differ from the characteristics of another kind of matter is called a substance.

Examples: Oxygen, water, carbon dioxide, urea, glucose, common salt etc are different substances.

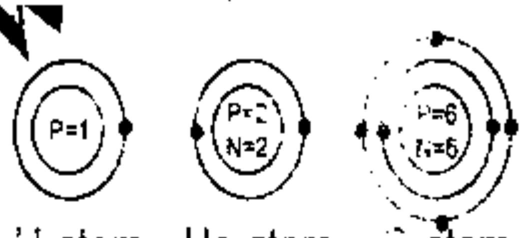
Element: An element is a substance whose all the atoms have the same atomic number.
A substance that cannot be converted to other simpler substances is called an element

Examples: Substances like oxygen, hydrogen, iron, copper, aluminium etc are elements.

Atomic Number: The number of protons in the nucleus of an atom is known as its atomic number.

Example: There is only one proton in the nucleus of H-atom, therefore, its atomic number is 1.

Note: All the atoms of a given element have the same number of protons and therefore the same atomic number.



H-atom He-atom C-atom

Mass Number: The total number of protons and neutrons in an atom is known as its mass number.
Number of neutrons = mass number – atomic number

b. When natural gas burns, CO_2 is formed. If 0.25 moles of CO_2 is formed, what mass of CO_2 is produced?

Problem solving strategy:

Carbon dioxide is a molecular substance. Determine its molar mass and use it to convert moles to mass in grams.

0.25 moles of CO_2 ---→ ? g of CO_2

Solution:

Molar mass of CO_2 = 12 + 16 × 2

Q6. a. Define Avogadro's number. How does it relate to the masses of chemical substances?

Ans: Avogadro's number:

The number of representative particles in one mole of the substance is known as Avogadro's number

The Avogadro's constant or number 6.02×10^{23} is the number of atoms of carbon in exactly 12 g of carbon-12.

Relation of Avogadro's number to the masses of chemical substances:

One mole of hydrogen atoms = 6.02×10^{23} atoms of hydrogen (H) = 1 a.m.u

One mole of hydrogen molecules = 6.02×10^{23} molecules of hydrogen (H_2) = 2 a.m.u

One mole of water = 6.02×10^{23} molecules of water (H_2O) = 18 a.m.u

b. How many moles of each of the following substance are present?

(a) A balloon filled with 5g of hydrogen. (b) A block of ice that weighs 100g

Problem solving strategy:

Hydrogen and ice both are molecular substances. Determine their molar masses. Use the molar mass of each to convert masses in grams to moles.

mass $\longrightarrow$? moles

Solution:

a) Molar mass of H_2 $= 1.008 \times 2 = 2.016g$

1 mole of H_2 $= 2.016g$

So, $2.016g$ of H_2 = 1 mole of H_2

$1g$ of H_2 $= \frac{1}{2.016}$ moles of H_2

$5g$ of H_2 $= \frac{1}{2.016} \times 5$ moles of H_2 = 2.48 moles of H_2O

b) 1 mole of ice (H_2O) $= 2 \times 1.008 + 16 = 2.016 + 16 = 18.016$ g

1 mole of H_2O = 18.016 g

So, 18.016g of H_2O = 1 mole

1g of H_2O $= \frac{1}{18.016}$ moles

100g of H_2O $= \frac{1}{8.016} \times 100$ moles = 5.55 moles of H_2O

IMPORTANT SHORT QUESTION ANSWERS
(Reduced Syllabus)

Q1. How is mole related to Avogadro's number?

Ans: Relation of Avogadro's number to mole:

One mole of hydrogen atoms = 6.02×10^{23} atoms of hydrogen (H) = 1 a.m.u

One mole of hydrogen molecules = 6.02×10^{23} molecules of hydrogen (H_2) = 2 a.m.u

One mole of water = 6.02×10^{23} molecules of water (H_2O) = 18 a.m.u

$$\text{Number of moles of a substance} = \frac{\text{Number of molecules of substance}}{N_A}$$

Number of moles of a substance = Number of molecules of substance

SELF ASSESSMENT EXERCISE 1.9

1. The molecular formula of a compound used for bleaching hair is H_2O_2. Calculate
   (a) Mass of this compound that would contain 2.5 moles.
   (b) Number of moles of this compound that would exactly weigh 30 g.

Solution: Number of moles = 2.5 moles
Molar mass of hydrogen peroxides (H_2O_2) = 2 × 1 + 2 × 16 = 2+ 32 = 34 g
Mass of H_2O_2 = number of moles × Molar mass = 2.5 × 34 = 85 g

b. Mass of hydrogen peroxides (H_2O_2) = 30 g
Molar mass of hydrogen peroxides (H_2O_2) = 2 × 1 + 2 × 16 = 2+ 32 = 34 g
Number of moles = ?

$$\text{Number of moles} = \frac{\text{Mass in grams}}{\text{Molar mass}}$$

$$\text{Number of moles} = \frac{30}{34} = 0\ 88 \text{ moles}$$

2. A spoon of table salt, NaCl contains 12.5 grams of this salt. Calculate the number of moles it contains.

Solution: Mass of sodium chloride (NaCl) = 12.5 g
Molar mass of sodium chloride (NaCl) = 23 + 35.5 = 58.5 g
Number of moles = ?

$$\text{Number of moles} = \frac{\text{Mass in grams}}{\text{Molar mass}}$$

$$\text{Number of moles} = \frac{12.5}{58.5} = 0.21 \text{ moles}$$

3. Before the digestive systems X-rayed, people are required to swallow suspensions of barium sulphate ($BaSO_4$). Calculate mass of one mole of $BaSO_4$.

Solution: Mass of 1 mole of barium sulphate ($BaSO_4$) = 137 + 32 + 4 × 16 = 137 + 32 + 64 = 233 g

SELF ASSESSMENT EXERCISE 1.10

1. Aspirin is a compound that contains carbon, hydrogen and oxygen. It is used as a painkiller. An aspirin tablet contains 1.25×10^{30} molecules. How many moles of this compound are present in the tablet?

Solution: Number of molecules = 1.25×10^{30} molecules
Avogadro's number = $N_A = 6.022 \times 10^{23}$
Number of moles = ?

$$\text{Number of moles} = \frac{\text{Number of molecules}}{N_A}$$

$$\text{Number of moles} = \frac{1.25 \times 10^{30}}{6.022 \times 10^{23}} = 0.2076 \times 10^{30-23} = 0.2076 \times 10^{7}$$

Number of moles = 2.076×10^{6} moles

2. A method used to prevent rusting in ships and underground pipelines involves connecting the iron to a block of a more active metal such as magnesium. This method is called cathodic protection. How many moles of magnesium are present in 1 billion (1×10^9) atoms of magnesium.

Solution: Number of atoms = 1×10^9 atoms
Avogadro's number = $N_A = 6.022 \times 10^{23}$

Number of moles = $\frac{1 \times 10^9}{6.022 \times 10^{23}} = 0.166 \times 10^{9-23} = 0.166 \times 10^{-14}$

Number of moles = 1.66×10^{-15} moles

Q2. TNT or trinitrotoluene is an explosive compound used in bombs. It contains 7 C-atoms, 6 H-atoms, 5 N-atoms and 6 O-atoms. Write its empirical formula.

Ans: $C_7H_6N_5O_6$

Q3. A molecule contains four phosphorus atoms and ten oxygen atoms. Write the empirical formula of this compound. Also determine the molar mass of this molecule.

Ans: Molecular formula = P_4O_{10} , Empirical formula = P_2O_5.

Molar mass of P_4O_{10} = 4 × 31 +10 × 16 = 124 + 160 = 284 g

Q4. Indigo ($C_{16}H_{10}N_2O_2$), the dye used to colour blue jeans is derived from a compound known as indoxyl (C_8H_7ON). Calculate the molar masses of these compounds. Also write their empirical formulas.

Ans: Molar mass of Indigo ($C_{16}H_{10}N_2O_2$) = 16 × 12 + 10 ×1 + 2 × 14 + 2×16 = 192 + 10 + 28 + 32 = 262 g

Empirical formula of Indigo ($C_{16}H_{10}N_2O_2$) = C_8H_5NO (16 : 10 : 2 : 2)

Molar mass of Indoxyl (C_8H_7ON) = 8 × 12 + 7 ×1 + 16 +14 = 96 + 7+16 +14 = 133 g

Empirical formula of Indoxyl (C_8H_7ON) = C_8H_7ON (8 : 7 : 1 : 1)

Q5. Identify the substance that has formula mass of 133.5 amu.

a. $MgCl_2$ b. S_2Cl_2 c. BCl_3 d. $AlCl_3$

Solution:

a. Formula mass of $MgCl_2$ = 24 + 2 × 35.5 = 24 + 71 = 95 amu

b. Formula mass of S_2Cl_2 = 2 × 32 + 2 × 35.5 = 64 +71 = 135 amu

c. Formula mass of BCl_3 = 11 + 3 × 35.5 = 11 + 106.5 = 117.5 amu

d. Formula mass of $AlCl_3$ = 27 + 3 × 35.5 = 27 + 106.5 = 133.5 amu

Hence Formula mass of $AlCl_3$ is 133.5 amu therefore option 'd' is correct.

Q6. What mass of sodium metal contains the same number of atoms as 12.00g of carbon?

Solution: Avogadro s number = $N_A = 6.022 \times 10^{23}$

1 mole of sodium atom = 1 mole of carbon -12 = 6.022×10^{23} atoms.

Therefore 23 g of sodium metal contains the same number of atoms as 12.00g of carbon.

Q7. What mass of oxygen contains the same number of molecules as 42g of nitrogen?

Solution: 1 mole of N = 14 g

3 mole of N = 3 x 14 = 42 g

Similarly, 1 mole of O = 16 g

3 mole of O = 3 x 16 = 48 g

Q8. Calculate the mass of one hydrogen atom in grams.

Ans: Mass of 1 hydrogen atom = $\frac{\text{Atomic mass of hydrogen}}{\text{Avogadro's number}}$

Mass of 1 hydrogen atom = $\frac{1.008}{6.022 \times 10^{23}} = 0.1674 \times 10^{-23} = 1.674 \times 10^{-24}$ g

Q9. Calculate the number of H-atoms present in 18g H_2O.

Ans: Known mass of H_2O = 18 g = 1 mole of H_2O = 6.022×10^{23} atoms.

Since one molecule of water has two atoms of hydrogen, therefore,

Number of hydrogen atoms in 18 g of H_2O = $2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$ atoms.

GUESS PAPER & MODEL PAPER # 2
BASED ON CHAPTER # 2 (Reduced Syllabus)
STRUCTURE OF ATOMS

CHAPTER 2 : STRUCTURE OF ATOM

Rutherford atomic model (experiment, Defects), Bohr's atomic theory, electronic configuration

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) According to Bohr atomic model:
A Each orbit has fixed energy, so each orbit is called sub-energy level
B. The energy of the electron is inversely proportional to its distance from the nucleus.
C Light is absorbed when an electron jumps a lower energy orbit.
D. The farther the electron is from the nucleus, the more energy it has.

(ii) s sub-shell can accommodate maximum ________ electrons.
A 2 B 6 C 10 D. 14

(iii) p sub-shell can accommodate maximum ________ electrons.
A 2 B. 6 C 10 D. 14

(iv) f sub-shell can accommodate maximum ________ electrons.
A. 2 B. 6 C. 10 D. 14

(v) M shell has sub-shells:
A. 1s, 2s B. 2s. 2p C 3s, 3p, 3d D 1s, 2s, 3s

(vi) A sub-shell that can accommodate 6 electrons is
A. s B d C. p D. f

(vii) $_{11}Na$ has electronic configuration:
A $1s^2 2s^2 3s^1$ B $1s^2 2s^2 2p^7$ C. $1s^2 2s^2 2p^5 3s^2$ D. $1s^2 2s^2 2p^6 3s^1$

(viii) Rutherford used _____ particles in his experiments.
A He atoms B. He^+ C. He^{+2} D. He^{-2}

(ix) d sub-shell can accommodate maximum ________ electrons.
A 2 B 6 C. 10 D 14

(x) Rutherford bombarded a very thin gold foil about ________ thickness with α-particles.
A. 0.0004cm B. 4cm C. 0.4cm D. 0.04cm

(xi) An element has six electrons in M-shell, its atomic number is:-
A. 14 B. 16 C 18 D. 8

(xii) The value of h is:-
A. 6.625×10^{-36} Js B. 6.626×10^{-34} Js

CHEMISTRY SSC-I

Time allowed: 2:40 hours Total Marks: 53

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet-B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. Distinguish between shell and sub-shell.
ii. An atom is electrically neutral, why?
iii. How many sub-shells are there in N shell.
iv. Give notation for sub-shells of M shell.
v. List the sub-shells of M Shell in order of increasing energy.
vi. What are the defects of Rutherford's model of atom?
vii. How did Bohr's atomic theory modified Rutherford's model of atom with the help of Quantum theory? Explain the postulates of Bohr's atomic theory?
viii. What do you mean by shells/orbits and energy level?

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. Draw Bohr's Model for the following atoms indicating the location for electron, protons and neutrons:
  a. Potassium (Atomic No 19, Mass No. 39)
  b. Silicon (Atomic No. 14 Mass No. 28)
  c. Argon (Atomic No. 18 Mass No. 40)
ii. Write electronic configuration for the following elements:
iii. Define Auf Bau principle.
iv. What do you mean by electronic configuration. Explain by taking specific examples?
v. The atomic number of an element is 23 and its mass number is 56.
  How many protons and electrons does an atom of this element have?
vi. The atomic number of an element is 23 and its mass number is 56.
  How many neutrons does this atom have?
vii. Describe the accommodation of maximum electrons in s sub-shell, p sub-shell, d sub-shell and f sub-shell.

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 a. The atomic symbol of aluminum is written as $^{27}_{13}Al$. What information do you get from it?
b. Describe the contribution that Rutherford made to the development of the atomic theory.

Q.5 a. Explain how Bohr's atomic theory differed from Rutherford's atomic theory.
b. Describe the presence of sub shells in a shell.

Q.6 a. How did Rutherford's model of an atom first of all proved the existence of nucleus in an atom? What are the conclusions drawn by Rutherford from the scattering experiment of α - particles?
b. The atomic number of an element is 23 and its mass number is 56.
  i. How many protons and electrons does an atom of this element have?
  ii. How many neutrons does this atom have?

SOLUTION OF GUESS PAPER & MODEL PAPER # 2
(Reduced Syllabus)

SECTION- A (MCQs)

i. D	ii. A	iii. B	iv. D	v. C	vi. C
vii. D	viii. C	ix. C	x. A	xi. B	xii. B

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. **(6 × 3 = 18)**

i. Distinguish between shell and sub-shell.

Ans: Shell:

An electron shell may be thought of as an orbit followed by electrons around an atom's nucleus. The closest shell to the nucleus is called the "1 shell" (also called "K shell"), followed by the "2 shell" (or "L shell"), then the "3 shell" (or "M shell"), and so on further and further from the nucleus. The shell letters K, L, M, ... are alphabetical.

Each shell can contain only a fixed number of electrons: The 1 shell can hold up to two electrons, the 2 shell can hold up to eight electrons, and in general, the n shell can hold up to $2n^2$ electrons

Shell the highest way of classing an electron, contains many subshells. Labelled with 1, 2, 3, 4, 5 etc.

Sub-shell:

Each shell is composed of one or more subshells, which are themselves composed of atomic orbrtals.

For example, the first (K) shell has one subshell, called "1s"; the second (L) shell has two subshells, called "2s" and "2p"; the third shell has "3s", "3p", and "3d"; and so on.

Sub-shell a group of orbitals with particular properties like shape and angular momentum. Labelled s, p, d, f, g etc.

ii. An atom is electrically neutral, why?

Ans: Protons have a positive charge that is usually expressed as **+1** though in terms of coulombs it is **$+1.602\times10^{-19}$ C**. The electron has a negative charge that is expressed as -1 and in terms of coulombs is equal to **-1.602×10^{-19} C.**

A neutral atom has the same number of electrons as the number of protons in it. Therefore the net charge in an atom is equal to zero. This makes an atom electrically neutral.

iii. How many sub-shells are there in N shell.

Ans: Shell **n = 4** contains 4 sub shells, **s, p, d and f (4s, 4p, 4d, 4f).**

iv. Give notation for sub-shells of M shell.

Ans: Notation for M shell is n =3 So M shell has 3 sub-shells called **3s , 3p and 3d.**

v. List the sub-shells of M Shell in order of increasing energy.

Ans: **3s < 3p < 3d**

vi. What are the defects of Rutherford's model of atom?

Ans: Defects in Rutherford's atomic model:

Rutherford's model of an atom resembles our solar system. It has following defects:

1. Classical physics suggests that electron being charged particle will emit energy continuously while revolving around the nucleus. Thus the orbit of the revolving electron becomes smaller and smaller until it would fall into the nucleus. This would collapse the atomic structure.
2. If revolving electron emits energy continuously it should form a continuous spectrum for an atom but a line spectrum is obtained

Bohr formulated new explanation and a new theory to remove defects from the Rutherford's atomic model.

vii. How did Bohr's atomic theory modified Rutherford's model of atom with the help of Quantum theory? Explain the postulates of Bohr's atomic theory?

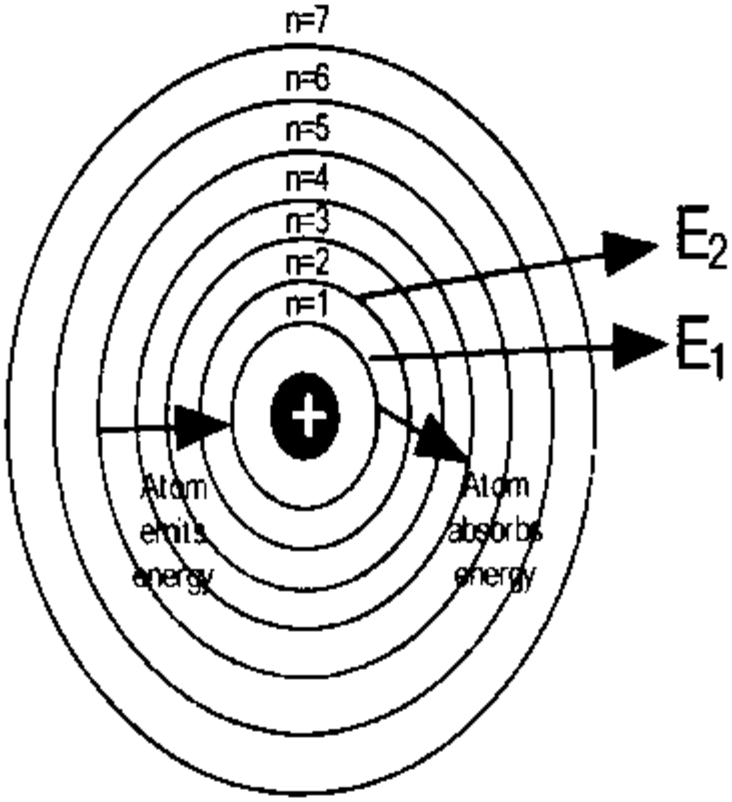
Ans: Bohr's atomic theory:

But it also explains the observed line spectrum of the hydrogen atom. Main postulates of Bohr's atomic theory are as follows:

1. The electron in an atom revolves around the nucleus in one of the circular orbits. Each orbit has a fixed energy. So each orbit is also called energy level.
2. The energy of the electron in an orbit is proportional to its distance from the nucleus. The farther the electron is from the nucleus, the more energy it has.
3. The electron revolves only in those orbits for which the angular momentum of the electron is an integral multiple of $\frac{h}{2\pi}$ where h is Plank's constant (its value is 6.626×10^{-34} J.s).
4. Light is absorbed when an electron jumps to a higher energy orbit and emitted when an electron falls into a lower energy orbit. Electron present in a particular orbit does not radiate energy.
5. The energy of the light emitted is exactly equal to the difference between the energies of the orbits

$$\Delta E = E_2 - E_1$$

Where ΔE is the energy difference between any two orbits with energies E_1 and E_2



Bohr's model of the atom

viii. What do you mean by shells/orbits and energy level?

Ans: Shells or Orbits:

According to Bohr's atomic theory, the electron in an atom revolves around the nucleus in one of the circular paths called shells or orbits.

Each shell is described by an n value. n can have values 1,2,3.....

When, **n = 1**, it is K shell
n = 2, it is L shell
n = 3, it is M shell etc.

As the value of n increases distance of electron from the nucleus and energy of the shell increases.

Energy level:

Each shell has a fixed energy. So each shell is also called energy level.

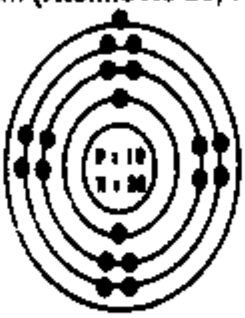
SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

c. Argon (Atomic No. 18 Mass No. 40)

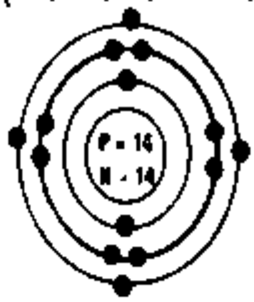
Ans:

a. Potassium (Atomic No 19, Mass No. 39)



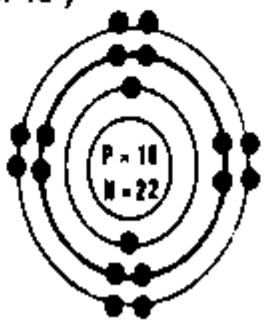
(K = 2, L = 8, M = 8, N = 1)

b. Silicon (Atomic No. 14 Mass No. 28)



(K = 2, L = 8, M = 4)

c. Argon (Atomic No. 18 Mass No. 40)



(K = 2, L = 8, M = 8)

ii. Write electronic configuration for the following elements:

Solution:

a.	$^{28}_{14}Si$	**Si**	**(atomic number 14):**	$1s^2\,2s^2\,2p^6\,3s^2\,3p^2$
b.	$^{24}_{12}Mg$	**Mg**	**(atomic number 12):**	$1s^2\,2s^2\,2p^6\,3s^2$
c.	$^{27}_{13}Al$	**Al**	**(atomic number 13):**	$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$
d.	$^{40}_{18}Ar$	**Ar**	**(atomic number 18):**	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$

iii. Define Auf Bau principle.

Ans: Auf Bau principle:

We can fill the electrons present in various elements by using Auf Bau Principle.

According to this principle, electrons fill the lowest energy sub-shell that is available first. This means electron will fill first 1s, then 2s, then 2p and so on.

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \ \ldots$$

iv. What do you mean by electronic configuration. Explain by taking specific examples?

Ans: Electronic configuration:

Electronic configuration is the distribution/arrangement of electrons among the different sub-shells of an atom.

Examples:

Electronic configuration of hydrogen:

Hydrogen has atomic number 1. So it has only one electron that will occupy lowest energy sub-shell 1s. The electronic configuration of H is $\mathbf{1s^1}$

Electronic configuration of helium:

Helium has atomic number 2, so it has two electrons. Since s sub-shell can accommodate two electrons, so electronic configuration of He is $\mathbf{1s^2}$.

Electronic configuration of lithium:

Lithium has atomic number 3, so it has three electrons, two will fill 1s sub-shell and one 2s sub-shell. So electronic configuration of Li is $\mathbf{1s^2 2s^1}$

$_5B = 1s^2 2s^2 sp^1$

$_6C = 1s^2 2s^2 2p^2$

$_7N = 1s^2 2s^2 2p^3$

$_8O = 1s^2 2s^2 2p^4$

$_9F = 1s^2 2s^2 2p^5$

$_{10}Ne = 1s^2 2s^2 2p^6$

$_{11}Na = 1s^2\ 2s^2\ 2p^6\ 3s^1$

$_{12}Mg = 1s^2\ 2s^2\ 2p^6\ 3s^2$

v. The atomic number of an element is 23 and its mass number is 56. How many protons and electrons does an atom of this element have?

Solution: Atomic number = 23 Atomic mass = 56

Atomic number = Z = Number of protons = 23

Atomic number = Z = Number of electrons = 23

vi. The atomic number of an element is 23 and its mass number is 56. How many neutrons does this atom have?

Solution: Number of neutrons = Atomic mass – Atomic Number = A – Z

Number of neutrons = 56 – 23 = 33

vii. Describe the accommodation of maximum electrons in s sub-shell, p sub-shell, d sub-shell and f sub-shell.

Ans: **Accommodation of electrons in sub-shells:**

s sub-shell can accommodate maximum 2 electrons.

p sub-shell can accommodate maximum 6 electrons.

d sub-shell can accommodate maximum 10 electrons.

f sub-shell can accommodate maximum 14 electrons.

The increasing order of energy of the sub-shells belonging to different shells is given below.

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \ldots$$

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 a. The atomic symbol of aluminum is written as $^{27}_{13}Al$. What information do you get from it?

Ans: Atomic number = 13, Atomic mass = 27, Number of electrons = 13,
Number of proton = 13, Number of neutron = 27 – 13 = 14.

Distribution of electrons in different shells: K = 2, L = 8, M = 3

Aluminum(Al) has 3 electrons in its valence shell. Aluminum can loss 3 electrons to form Al^{+3} therefore aluminum is a metal.

b. Describe the contribution that Rutherford made to the development of the atomic theory.

Ans: **Rutherford's Atomic Model:**

On the basis of conclusions drawn from his experiments, Rutherford proposed a new model for an atom. He proposed a planetary model (similar to the solar system) for an atom

i. An atom is neutral particle.

ii. The mass of an atom is concentrated in a very small dense positively charged region. He named this region as nucleus. A positively charged region is present at the centre of an atom.

iii. The electrons are revolving around the nucleus in circles. These circles are called orbits. The Centripetal force cue to the revolution of the electrons balances the electrostatic force of attraction between the nucleus and electron.

Explain how Bohr's atomic theory differed from Rutherford's atomic theory.

Bohr proposed that an electron moves around the nucleus in well defined circular paths called orbits. An orbit has fixed energy. Electron present in an orbit does not emit energy. Bohr atomic theory explains nicely the stability of an atom and also explains why an atom gives line spectrum.

b. Describe the presence of sub shells in a shell.

Ans: Sub-shells:

A shell or energy level is sub divided into sub-shells or sub-energy levels. n value of a shell is placed before the symbol for a sub-shell.

For instance:

i. **For K shell:** n = 1, for K shell. It has only one sub-shell which as represented by 1s.
ii. **For L shell:** n=2, L shell has two sub-shells. these are designated as 2s and 2p.
iii. **For M shell:** n =3 So M shell has 3 sub-shells called 3s , 3p and 3d.
iv. **For N shell:** N shell has 4s, 4p, 4d and 4f sub-shells.

Note: s sub-shell can accommodate maximum 2 electrons
p sub-shell can accommodate maximum 6 electrons.
d sub-shell can accommodate maximum 10 electrons.
f sub-shell can accommodate maximum 14 electrons.

The increasing order of energy of the sub-shells belonging to different shells is given below

1s < 2s < 2p < 3s < 3p< 4s< 3d .....

Q.6 How did Rutherford's model of an atom first of all proved the existence of nucleus in an atom? What are the conclusions drawn by Rutherford from the scattering experiment of α - particles?

Ans: Rutherford's atomic model:

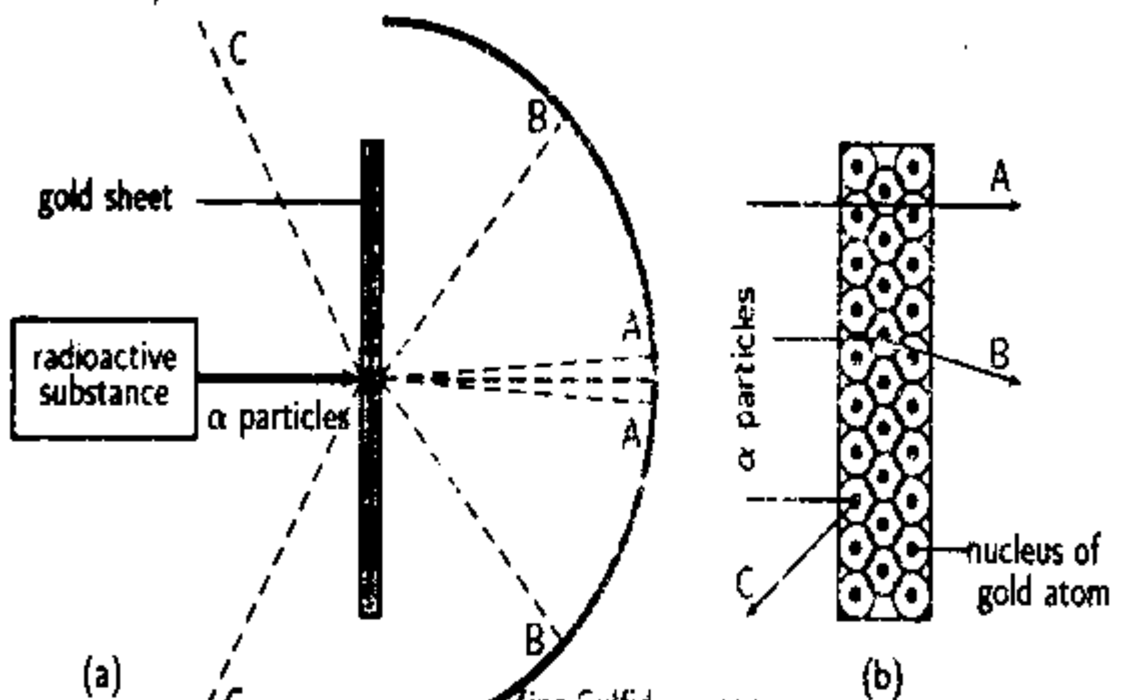
In 1911 Rutherford performed an experiment in order to know the arrangement of electrons and protons in atoms.

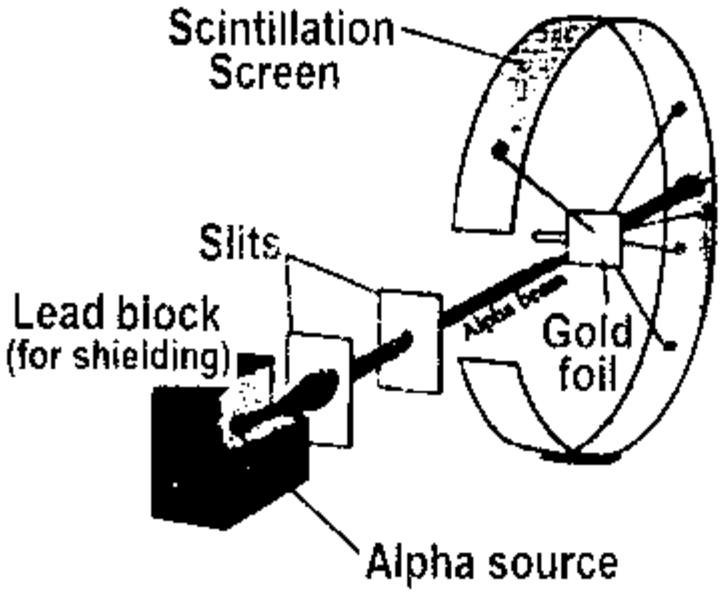
Rutherford's experiment:

Rutherford bombarded a very thin gold foil about 0.0004cm thickness with α −particles. He used α −particles obtained from the disintegration of polonium. α −particles are helium nuclei that are doubly positively charged (He^{++}).

Most of these particles passed straight through the foil. Only few particles were slightly deflected. But one in 1 million was deflected through an angle greater than 90° from their straight paths.

Rutherford performed a series of experiments using thin foils of other elements. He observed similar results from these experiments.





Rutherford's Experiment

Rutherford draw following conclusions:

1. Since majority of the α-particles passed through the foil undeflected most of the space occupied by an atom must be empty
2. The deflection of a few α-particles through angles greater than 90° shows that these particles are deflected by electrostatic repulsion between the positively charged α-particles and the positively charged part of atom
3. Massive α-particles are not deflected by electrons

Discovery of nucleolus:

Rutherford proposed a planetary model (similar to the solar system) for an atom. An atom is neutral particle. The mass of an atom is concentrated in a very small dense positively charged region. He named this region as nucleus

A positively charged region is present at the centre of an atom and the electrons are revolving around the nucleus in circles. These circles are called orbits

Note: The centripetal force due to the revolution of electrons balances the electrostatic force of attraction between the nucleus and electron

IMPORTANT SHORT QUESTION ANSWERS (Reduced Syllabus)

SELF ASSESSMENT EXERCISE 2.3

Write the complete electronic configuration for the following elements;

Al	(atomic number 13)	Si	(atomic number 14)
P	(atomic number 15)	S	(atomic number 16)
Cl	(atomic number 17)	Ar	(atomic number 18)

Solution:

Al (atomic number 13): $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^1$
Si (atomic number 14): $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^2$
P (atomic number 15): $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^3$
S (atomic number 16): $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^4$
Cl (atomic number 17): $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^5$

SELF ASSESSMENT EXERCISE 2.4

Write the electronic configuration for the following isotopes.

(a) $^{14}_{6}C$, **(b)** $^{35}_{17}Cl$, **(c)** $^{37}_{17}Cl$

Solution:

(a) $^{14}_{6}C$ $1s^2 2s^2 2p^2$ (b) $^{35}_{17}Cl$ $1s^2 2s^2 2p^6 3s^2 3p^5$ (c) $^{37}_{17}Cl$ $1s^2 2s^2 2p^6 3s^2 3p^5$

Q1. The electronic configurations listed are incorrect. Explain what mistake has been made in each and write correct electronic configurations.

$x = 1s^2 2s^2 2p^4 3p^2$
$y = 1s^2 2s^1 2p^1$
$z = 1s^2 2s^2 2p^5 3s^1$

Ans: According to Auf Bau principle, electrons fill the lowest energy sub-shell that is available first. This means electron will fill first 1s, then 2s, then 2p and so on

$1s < 2s < 2p < 3s < 3p < 4s < 3d$ .....

$x = 1s^2 2s^2 2p^4 3p^2$ Correct electronic configuration = $1s^2 2s^2 2p^6$
$y = 1s^2 2s^1 2p^1$ Correct electronic configuration = $1s^2 2s^2$
$z = 1s^2 2s^2 2p^5 3s^1$ Correct electronic configuration = $1s^2 2s^2 2p^6$

Q2. Which orbital in each of the following pairs is lower in energy?

a. 2s, 2p **b.** 3p, 2p **c.** 3s, 4s

Ans: **a.** 2s, 2p The energy of 2s < 2p **b.** 3p, 2p The energy of 2p < 3p.
c. 3s, 4s The energy of 3s < 4s.

Note: According to the Auf Bau principle the increasing order of energy of the sub-shells belonging to different shells is given below.

$1s < 2s < 2p < 3s < 3p < 4s < 3d$ .....

Q3 How testing prevailing theories bring about changes in them?

Ans: When ideas of scientists are not correct. Scientists did not discard his theory. Instead, they revised the theory to take into account new discoveries. This shows how testing prevailing theories bring about changes in them

Q4. How experimental results of some scientists help chemist to formulate new theories and new explanation.

Ans: Definition To "Result" = Something that results-effect, consequence-beneficial or discernible effect-something obtained by calculation or investigation

Definition To "Experimental" = A controlled procedure carried out discover or test something.

Example: You water a plant to see if it would grow during the week. The experiment is to see if the plant would grow during the week, and at the end of the week you find out that the plant did grow. So the result would be that it did grow during the week you watered it

Example: Bohr atomic theory explains nicely the stability of an atom and also explains why an atom gives line spectrum. Development of Bohr's atomic model explains how interpretations of experimental results of other scientists help chemists to formulate new explanations and new theories.

Q5. How many electrons can be placed in all of the sub-Shells in the n=2 shell?

Ans: When n = 2, it is L-shell. L-shell can accommodate electrons = $2n^2 = 2(2)^2 = 8$.
L-shell has 2 sub-shell's i.e. s and p, distribution of electrons in L-shell is $2s^2\ 2p^2$

GUESS PAPER & MODEL PAPER # 3
BASED ON CHAPTER # 3 (Reduced Syllabus)
PERIODICITY OF PROPERTIES

CHAPTER 3 : PERIODIC TABLE AND PERIODICITY OF PROPERTIES

Periods and groups of element, s and p blocks in periodic table, Periodicity of properties, Shielding effect and trends, atomic size and trends, Ionization energy and trends, Electron affinity and trends, electronegativity and trends.

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) Number of periods in the periodic table are:
A. 8 B. 7 C. 16 D. 5

(ii) Which of the following groups contain alkaline earth metals?
A 1A B IIA C VIIA D. VIIIA

(iii) Which of the following elements belongs to VIIIA?
A. Na B Mg C. Br D. Xe

(iv) Main group elements are arranged in ________ groups.
A. 6 B 7 C. 8 D. 10

(v) Period number of $^{27}_{13}Al$ is:
A. 1 B. 2 C. 3 D. 4

(vi) Valence shell electronic configuration of an element M (atomic no. 14) is:
A $2s^22p^1$ B. $2s^22p^2$ C $2s^22p^3$ D. $4s^1$

(vii) Which of the following elements you expect to have greater shielding effect?
A Li B. Na C. K D. Rb

(viii) As you move from right to left across a period, which of the following do not increase:
A. electron affinity B. ionization energy
C. nuclear charge D. shielding effect

(ix) All the elements of Group IIA are less reactive than alkali metals. This is because these elements have:
A. low ionization energies B relatively greater atomic sizes
C similar electronic configuration D. decreased nuclear charge

(x) The shielding effect in $_{11}Na$ is:
A. greater than that of $_3Li$ B greater than that of $_{19}K$
C. equal to that of $_{19}K$ D. equal to that of $_3Li$

(xi) Hydrogen after losing one electron forms H^+. In this property, it resembles:
A. transition elements B. alkaline earths
C. halogens D. alkali metals

(xii) The atomic radii of the elements in Periodic Table:

CHEMISTRY SSC-I

Time allowed: 2:40 hours Total Marks: 53

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet–B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. Which atom has the smaller ionization energy?
(a) B or N (b) Be or Mg (c) C or Si

ii. Which atom has higher shielding effect, Li or Na?

iii. Explain why, Na has higher ionization energy than K?

iv. Alkali metals belong to s-block in the periodic table, why?

v. Arrange the elements in each of the following groups in order of increasing ionization energy: (a) Li, Na, K (b) Cl, Br, I

vi. Arrange the elements in each of the following in order of decreasing shielding effect.
(a) Li, Na, K (b) Cl, Br, I (c) Cl, Br

vii. Specify which of the following elements you would expect to have the greatest electron affinity. S, P, Cl

viii. Write electron dot symbols for an atom of the following elements
(a) Be (b) K (c) N (d) I

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. In what region of the periodic table you will find elements with relatively
a) high ionization energies b) low ionization energies

ii. Choose the element whose atom you expect to have smaller atomic radius in each of the following pairs. (a) O or S (b) O or F

iii. What do you mean by group or family in the periodic table?

iv. What are the representative and transition elements?

v. Choose the element whose atoms you expect to have smaller shielding effect.
(a) F or Cl (b) Li or Na (c) B or Al

vi. Which atom has greater shielding effect, Li or Na?

vii. Which atom has higher ionization energy, Li or Be?

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 a. What is meant by the periodicity of properties?

b. How does the shielding effect influence the atomic size, ionization energy and electron affinity? Justify that shielding effect increases down the group, and remains constant in a period from left to the right.

Q.5 a. What are the factors which influence the atomic size? How atomic radii vary in the periodic table?

b. Define electron affinity. Is it a energy releasing process? Give its variation in groups and periods.

Q.6 a. Define electronegativity. Name the most electronegative element and discuss its variation in groups and periods.

b. Define ionization energy. Highlight those factors which control the ionization energy values of elements in groups and periods.

SOLUTION OF GUESS PAPER & MODEL PAPER # 3
(Reduced Syllabus)

SECTION- A (MCQs)

i. B	ii. B	iii. D	iv. C	v. C	vi. B
vii. D	viii. D	ix. B	x. A	xi. D	xii. B

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. Which atom has the smaller ionization energy?
(a) B or N (b) Be or Mg (c) C or Si

Solution:

(a) B or N
Ionization energy increases across a period. The element that has smaller ionization energy will be further to the left in the periodic table. Therefore B has smaller ionization energy as compare to N. (Period 2)

(b) Be or Mg
Ionization energy decreases from top to bottom in a group. The element that has smaller ionization energy will correspond to the element closer to the bottom.
Therefore Mg has smaller ionization energy as compare to Be. (Group IIA)

(c) C or Si
Ionization energy decreases from top to bottom in a group. The element that has smaller ionization energy will correspond to the element closer to the bottom.
Therefore Si has smaller ionization energy as compare to C. (Group IVA)

ii. Which atom has higher shielding effect, Li or Na?

Ans: The valence-shell electron of $_{11}Na$ experience less attraction from the nucleus due to the presence of 10 inner-shell electrons as compares to $_3Li$ having 2 inner shell electrons.
Na atoms will have greater shielding effect due to greater number of inner shell electrons as compare to Li.

iii. Explain why, Na has higher ionization energy than K?

Ans: Ionization energy decreases from top to bottom in a group. The size of sodium (3 shells) is smaller than potassium (4 shells)
Therefore Na has higher ionization energy than K (group IA elements)

iv. Alkali metals belong to s-block in the periodic table, why?

Ans: s-Block Elements:
Groups IA on the left side of the table constitute s-Block because outer shell valence electrons of these elements are present in s-orbitals/sub shell.

v. Arrange the elements in each of the following groups in order of increasing ionization energy: (a) Li, Na, K (b) Cl, Br, I

Ans: (a) Li, Na, K
The ionization energy value decreases from top to bottom in a group. This is because the shielding effect in atoms increases as you descend. Greater shielding effects results in a weaker attraction of the nucleus for the valence electrons. So they are easier to remove. This leads to decrease in ionization energy from top to bottom in a group. Therefore increasing order of ionization energy is

Ionization energy of Li > Ionization energy of Na > Ionization energy of K

vi. Arrange the elements in each of the following in order of decreasing shielding effect.
(a) Li, Na, K (b) Cl, Br, I (c) Cl, Br

Ans: (a) Li, Na, K

As we move from top to bottom in a group the number of electronic shells increase. So the number of electrons in the inner shell also increase. As a result shielding effect increases.
Shielding effect of Li < shielding effect of Na < shielding effect of K
(Group IA elements)

(b) Cl, Br, I

As we move from top to bottom in a group the number of electronic shells increase So the number of electrons in the inner shell also increase. As a result shielding effect increases.
Shielding effect of Cl < shielding effect of Br < shielding effect of I
(Group VIIA elements)

(c) Cl, Br

As we move from top to bottom in a group the number of electronic shells increase. So the number of electrons in the inner shell also increase. As a result shielding effect increases.
Shielding effect of Cl < shielding effect of Br
(Group VIIA elements)

vii. Specify which of the following elements you would expect to have the greatest electron affinity. S, P, Cl

Ans: Variation in a period:

As we move from left to right across a period, the electron affinity generally increases This is due to increase in nuclear charge and decrease in atomic radius, which binds the extra electron more tightly to the nucleus Therefore Cl has greatest electron affinity as compare to S and P.

viii. Write electron dot symbols for an atom of the following elements
(a) Be (b) K (c) N (d) I

Ans:

(a) Be	(b) K	(c) N	(d) I
·Be·	K·	·N: (with dots)	:I: (with dots)

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. In what region of the periodic table you will find elements with relatively
a) high ionization energies b) low ionization energies

Ans: a) high ionization energies:

Noble gases (group VIIIA) have the highest ionization energies because they have complete outer most shells (follows octet or duplet rule). therefore It is difficult to remove an electron from their outer most shell.

Therefore, elements in the upper right of the periodic table have the highest ionization energy. (Noble gases, p-block).

b) low ionization energies:

On the other hand Alkali metals (group IA) have greater size, therefore Alkali metals have lowest ionization energy

Therefore, elements in the upper left of the periodic table have the lowest ionization energy (Alkalis metals , s-block)

ii. Choose the element whose atom you expect to have smaller atomic radius in each of the following pairs.
(a) O or S (b) O or F

Solution: (a) O or S:

iii. What do you mean by group or family in the periodic table?

Ans: Group or family:

Elements that have similar properties lie in the same column in the periodic table.

Each vertical column of elements in the periodic table is called a group or family.

iv. What are the representative and transition elements?

Ans. Representative elements/Normal elements:

Each group is identified by a number and the letter A or B. Group A elements are called **normal** or **representative elements**. They are also called main group elements.

All s-block and p-block elements excluding the noble gases are called **representative** or main group **elements**. These elements are also called **normal or typical elements**. Thus, on the basis of their location in the periodic table, the elements belonging to the group 1, 2 and 13 to 17 are the representative elements

Transition elements: Group B elements are called **transition elements**.

The d-block elements are called transition elements. These elements occupy group 3 to 12 of the periodic table and lie in between the s- and p-block elements. There are four series of d-block transition elements depending upon the energy level of the subshell which is in the process of filling in their outermost shell.

Inner-transition elements:

The f-block elements are called inner-transition elements. There are two series of the f-block elements. These are called 4f-series or lanthanides and 5f-series actinides

Outer Transition Elements: d-Block contains "Outer Transition Elements"

Inner Transition Elements: f-Block contains Inner Transition Elements

v. Choose the element whose atoms you expect to have smaller shielding effect.

(a) F or Cl (b) Li or Na (c) B or Al

Solution: (a) F or Cl

F atoms will have smaller shielding effect due to lesser number of inner shell electrons as compare to Cl

(b) Li or Na

Li atoms will have smaller shielding effect due to lesser number of inner shell electrons as compare to Na

(c) B or Al

B atoms will have smaller shielding effect due to lesser number of inner shell electrons as compare to Al

vi. Which atom has greater shielding effect, Li or Na?

Ans: The valence-shell electron of $_{11}$**Na** experience less attraction from the nucleus due to the presence of 10 inner-shell electrons as compares to $_3$Li having 2 inner shell-electrons.

Na atoms will have greater shielding effect due to greater number of inner shell electrons as compare to Li

vii. Which atom has higher ionization energy, Li or Be?

Ans: Ionization energy increases across a period. The element that has smaller ionization energy will be further to the left in the periodic table. Therefore Be has higher ionization energy as compare to Li

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. **($2 \times 10 = 20$)**

Q.4 a. What is meant by the periodicity of properties?

Ans: Periodicity of properties:

The electronic configuration of the elements show a periodic variation with the increasing atomic number. Therefore, the elements also show periodic variation in their physical and chemical properties. Elements having similar valence shell electronic configuration have been placed in the same group, one below the other.

Periodicity in Chemical properties:

Chemical properties depend on the valence shell electronic configuration. Because all the elements of a particular group have similar valence shell electronic configuration, they possess similar chemical characteristics.

In a period of periodic table the number of electrons present in the valance shell increase gradually left to right
chemical and physical properties also show variation in the same manner

b. **How does the shielding effect influence the atomic size, ionization energy and electron affinity? Justify that shielding effect increases down the group, and remains constant in a period from left to the right.**

Ans: Shielding Effect (Screening effect):

The reduction in force of attraction between nucleus and the valence electrons by the electrons present in the inner sub-shells is called shielding effect

OR

The decrease in the attractive force exerted by the nucleus on the valence-shell electrons due to the presence of the electrons lying between the nucleus and valence-shell, is called shielding effect. In other words, the **intervening electrons shield the valence-shell electrons from the nucleus.**

An important factor which affects the magnitude of shielding effect is the number of inner-shell electrons

Greater the number of inner-shell electrons or inner shells greater is the magnitude of shielding effect

Variation of shielding effect in a group:

As we move from top to bottom in a group the number of electronic shells increase So the number of electrons in the inner shell also increase As a result shielding effect increases

Variation of shielding effect in a period:

When we travel from left to the right in the periodic table in a period then number of shells remain the same So the shielding effect does not change in a period

Influence of shielding effect on atomic size:

Greater is the atomic size greater will be the shielding effect As the force of attraction Between the nucleus and outer electron decreases with the increases in distance The electron can be removed more easily or with less energy Moreover the force of attraction also decreases with the increasing shielding effect of the intervening electron

Influence of shielding effect on ionization energy:

Shielding effect depends upon the inner shell electrons If the inner shell electrons are greater shielding effect will be high and force of attraction between valence shell and nucleus will be low In a group, number of shells increases from top to bottom and shielding effect also increases The atoms with high shielding effect have low ionization energies

Influence of shielding effect on electron affinity:

In **groups**, the atomic radius increases with the increase in the proton number due to successive increase of electronic shell which also exert a **shielding effect** on the force of attraction between the nucleus and the valence electrons Thus the electron affinities usually **decrease** from top to bottom There are, of course **exceptions** to this generalization e g fluorine has electron affinity less than that of chlorine because it has a smaller atomic size than that of chlorine

Q.5 a. What are the factors which influence the atomic size? How atomic radii vary in the periodic table?

Ans: Atomic size/Atomic radius:

The size of an atom is the average distance between the nucleus of an atom and the outer electronic shell

Variation of atomic radii in the periodic table:

Variation of atomic radii in the period:

The atomic radius decreases in any given period as we move across the period This is because as we move from one element to the next on its right in a period Another electron is added to the same valence shell At the same time positive charge on the nucleus also increases by 1 The attractive force of the nucleus for the valence shell electron increases Therefore the shell size and atomic radius decreases

Example:

For example in going from lithium to beryllium atomic size decreases This we can understand from

Variation of atomic radii in the group:

The atomic radius increases in any given main group as we move down the group of elements. This is because the size of an atom is determined by the size of its valence shell. As we move to the next lower element in the group, the atom has an additional shell of electrons. This increases atomic radius.

Example:

For example, in going from **Li** to **Na** atomic radius increases. Consider electronic configuration of Li **($1s^2\ 2s^1$)** and **Na($1s^2$, $2s^2$, $2p^6$, $3s^1$)**. A new electronic shell has been added that increases atomic size.

b. Define electron affinity. Is it a energy releasing process? Give its variation in groups and periods.

Ans: Electron Affinity:

Electron affinity explains the anion formation. Electron affinity is defined as the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uninegative gaseous ion.

$$X(g) + e^- \longrightarrow X^-(g) + \text{electron affinity}$$

Electron affinity is an energy releasing process:

For example, for chlorine (Cl) atom, 349 kJ mol^{-1} of energy is released. This is the electron affinity of chlorine (Cl) and the process is **exothermic**.

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} + 349\ kj\ mol^{-1}$$

Variation of electron affinity in the periodic table:

(i) Variation of electron affinity in a group:

The electron affinity decreases from top to bottom in a group. This is due to increase in shielding effect. Due to increase in shielding effect added electron binds less tightly to the nucleus. As a result less energy is released.

(ii) Variation in a period:

As we move from left to right across a period, the electron affinity generally increases. This is due to increase in nuclear charge and decrease in atomic radius, which binds the extra electron more tightly to the nucleus. But shielding effect remains constant in each period. Therefore, alkali metals have lowest and halogens have the highest electron affinities in each period

Q.6 a. Define electronegativity. Name the most electronegative element and discuss its variation in groups and periods.

Ans: Electronegativity:

Electronegativity is the ability of an atom to attract the electrons towards itself in a chemical bond. The American chemist Linus Pauling devised a method for calculating relative electronegativities of elements

Most electronegative element:

Pauling assigned maximum value of 4.0 to the electronegativity of fluorine **($_9F$)**,

The least electronegative element is cesium (Cs) and is electronegative value is 0.7.

Variation of electronegativity in the periodic table:

(i) Variation in a group:

Electronegativity of elements decreases from top to bottom in a group.

This is because an increase in the atomic size decreases the tendency to attract the shared pair of electrons.

(ii) Variation in a period:

The electronegativity increases from left to right in a period. The nuclear charge increases from left to right while the electrons enter the same shell. The electrons in the same shell cannot shield each other effectively from the attractive force of the nucleus. Hence, the increased nuclear charge attracts the shared pair of electrons more strongly. This results in higher electronegativity.

Note: When ($X_A - X_B$) = 1.7, then A − B bond is 50 % ionic 50 % covalent.

When ($X_A - X_B$) < 1.7, then A − B bond is predominantly covalent.

H 2.1							He
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne 2.1
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.2	Ar 3.0
K 0.5	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 2.1
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9						

The electronegativities of elements.

b. Define ionization energy. Highlight those factors which control the ionization energy values of elements in groups and periods.

Ans: Ionization energy:

Ionization energy is defined as the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom

Example:

Ionization energy is a measure of the extent to which the nucleus attracts the outermost electron.

Explanation:

A high value of ionization energy means stronger attraction between the nucleus and the outermost electron.

Whereas a low ionization energy indicates a weaker force of attraction between the nucleus and the outermost electron.

Unit of ionization energy: Unit of ionization energy is $KJ/mole^{-1}$.

$$M_{(g)} + \text{ionization energy} \longrightarrow M^{+}_{(g)} + e^{-}$$

Variation of ionization energy in groups:

The ionization energy value decreases from top to bottom in a group. This is because the shielding effect in atoms increases as you descend. Greater shielding effects results in a weaker attraction of the nucleus for the valence electrons. So, they are easier to remove. This leads to decrease in ionization energy from top to bottom in a group.

Variation of Ionization energy in groups:

As we move from left to right in a period, the shielding effect remains constant. But progressively nuclear charge increases. A stronger force of attraction between nucleus and the valence electron increases. This leads to increase in ionization energy from left to right in a period

IMPORTANT SHORT QUESTION ANSWERS
(Reduced Syllabus)

SELF ASSESSMENT EXERCISE 3.1

In which period and group the following elements are present in the periodic table.

(a) Mg **(b)** Ne **(c)** Si **(d)** B

Solution:

(a) Mg

$$^{14}_{12}Mg = \underset{K}{\underline{1s^2}} \quad \underset{L}{\underline{2s^2, 2p^6}} \quad \underset{M}{\underline{3s^2}} \qquad \begin{matrix} K & L & M \\ 1 & 2 & 3 \end{matrix}$$

Valence shell is M

So n = 3, Mg is present in the 3^{rd} period.

Since total number of electrons in the valence shell are 2, it must be present in Group IIA.

(b) Ne

$$^{10}_{20}Ne = \underset{K}{\underline{1s^2}} \quad \underset{L}{\underline{2s^2, 2p^6}} \qquad \begin{matrix} K & L \\ 1 & 2 \end{matrix}$$

Valence shell is L

So n = 2, Ne is present in the 2^{nd} period.

Since total number of electrons in the valence shell are 2 + 6 = 8, it must be present in Group VIIIA.

(c) Si

$$^{14}_{28}Si = \underset{K}{\underline{1s^2}} \quad \underset{L}{\underline{2s^2, 2p^6}} \quad \underset{M}{\underline{3s^2, 3p^2}} \qquad \begin{matrix} K & L & M \\ 1 & 2 & 3 \end{matrix}$$

Valence shell is M

So n = 3, Si is present in the 3^{rd} period.

Since total number of electrons in the valence shell are 2 + 2 = 4, it must be present in Group IVA.

(d) B

$$^{9}_{5}B = \underset{K}{\underline{1s^2}} \quad \underset{L}{\underline{2s^2, 2p^1}} \qquad \begin{matrix} K & L \\ 1 & 2 \end{matrix}$$

Valence shell is L

So n = 2, B is present in the 2^{nd} period.

Since total number of electrons in the valence shell are 2+1=3, it must be present in Group IIIA.

GUESS PAPER & MODEL PAPER # 4
BASED ON CHAPTER # 4 (Reduced Syllabus)
STRUCTURE OF MOLECULES

CHAPTER 4: STRUCTURE OF MOLECULE

Chemical Bond, Types of bond, ionic bond, covalent bond, intermolecular forces.

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) Which of the following atoms will form an ion of charge -2 ?

	Atomic Number	Mass Number
A.	12	24
B.	14	28
C.	8	8
D	10	20

(ii) Which of the following atoms will not form cation or anion.
A. A (Atomic No. 16) B. B (Atomic No. 17)
C. C (Atomic No. 18) D. D (Atomic No. 19)

(iii) Which of the following atoms will form cation.
Atomic Number
A. 20 B. 18 C. 17 D. 15

(iv) Which of the following atoms obey duplet rule?
A. O_2 B. F_2 C. H_2 D. N_2

(v) Silicon belongs to Group IVA. It has ____ electrons in the valence shell
A. 2 B. 3 C. 4 D. 6

(vi) Phosphorus belongs to third period of Group VA. How many electrons it needs to complete its valence shell.
A. 2 B. 3 C. 4 D. 5

(vii) In the formation of AlF_3, aluminum atom loses ____ electrons.
A. 1 B. 2 C. 3 D. 4

(viii) Which of the following is not true about the formation of Na_2S:
A. Each sodium atom loses one electron B. Sodium forms cation
C. Sulphur forms anion D. Each sulphur atom gains one electron

(ix) Identify the covalent compound
A. NaCl B. MgO C. H_2O D. KF

(x) An ionic compound $A^+ B^-$ is most likely to be formed when:
A. the ionization energy of A is low and electron affinity of B is high

(xi) Helium ($_2$He) has only two electrons in its outermost K-shell called ________ rule.
A. octet B. triads C. duplet. D. octaves

(xii) An atom having six electrons in its valence shell will achieve noble gas electronic configuration by:
A. gaining one electron B losing all electrons
C. gaining two electrons D. losing two electrons

CHEMISTRY SSC-I

Time allowed: 2:40 hours Total Marks: 53

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet–B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)
i. Atoms combine to form various types of substances. But what holds them together?
ii. Explain formation of covalent bond between two nitrogen atoms.
iii. How does Al form cation?
iv. How does O form anion?
v. Write the names of two types of bond?
vi. Write one main difference between an ionic and a covalent bond.
vii. What is a polar covalent bond? Give two examples of the compounds containing polar covalent bond?
viii. Describe the formation of covalent bond between two non-metallic elements.

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)
i. What is the total number of shared electrons in a molecule of CO_2?
ii. Describe the factors affecting the formation of an ionic bond?
iii. What is meant by intermolecular force? Describe their types?
iv. What is the total number of shared electrons in a molecule of CO_2?
v. Discuss the nature and consequences of intermolecular forces?
vi. Define dipole-dipole interactions. What is the nature of these linkages?
vii. Describe the ways in which bonds may be formed.

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 a. How is an ionic bond formed? Explain with example.
b. What is a non polar covalent bond? Give two examples of the compounds containing non polar covalent bond?

Q.5 a. What is covalent bond. Describe the various types of covalent bonds. Give one example of each kind?
b. What is meant by hydrogen bonding? Discuss the nature and consequences of hydrogen bonding.

Q.6 a. For each of the following pairs of atoms, use electron dot and electron cross structures to write the equation for the formation of ionic compound.
(i) K and Cl (ii) Ca and S (iii) Al and N
b. Recognize the following compounds as having ionic bonds.

SOLUTION OF GUESS PAPER & MODEL PAPER # 4
(Reduced Syllabus)

SECTION- A (MCQs)

i. C	ii. C	iii. A	iv. C	v. C	vi. B
vii. C	viii. D	ix. C	x. A	xi. C	xii. C

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. $(6 \times 3 = 18)$

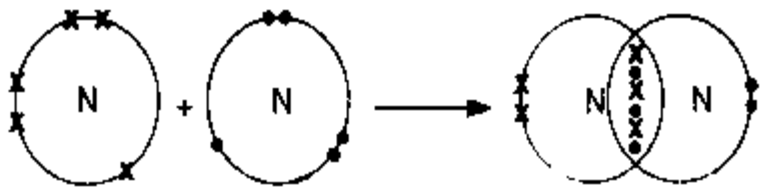
i. Atoms combine to form various types of substances. But what holds them together?

Ans: Fundamentally, some forces of attraction hold atoms together in substances. These forces are called chemical bonds. Basically the forces of attraction that lead to chemical bonding between atoms are electrical in nature. Electronic structure of an atom helps us to understand how atoms are held together to form substances. Atoms other than the noble gases have a tendency to react with other elements. These elements are reactive because they tend to gain stability by loosing or gaining electrons. When atoms gain or lose electron they acquire the configuration of next noble gas element.

When atoms share electrons they also acquire the configuration of next noble gas element.

ii. Explain formation of covalent bond between two nitrogen atoms.

Ans: Consider the formation of N_2 molecule. Nitrogen is in Group VA, so it has 5 electrons in the valence shell. It needs three electrons to complete its octet. So for sharing each N-atom contributes three electrons.



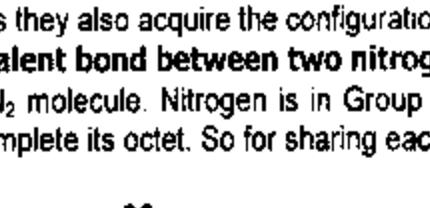
or $:N \equiv N:$

iii. How does Al form cation?

Ans: Formation of Al^{+3} ion:

Since Al atom has three electrons in the outer most shell. It losses three electrons to form Al^{+3} ion.

$$Al \quad 1s^2\, 2s^2 2p^6 3s^2 3p^1 \xrightarrow{-3e^-} Al^{+3} \quad 1s^2\, 2s^2 2p^6$$



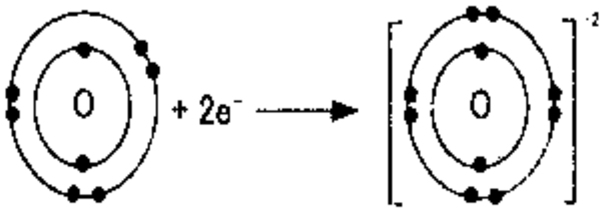
iv. How does O form anion?

Ans: Formation of anion by oxygen atom.

Oxygen belongs to Group VIA on the periodic table So it has six electrons in its valence shell. It needs two electrons to achieve noble gas configuration.

$$O \quad 1s^2 2s^2 2p^4 + 2e^- \longrightarrow O^{-2} \quad 1s^2\, \underbrace{2s^2 2p^6}_{\text{octet}}$$

You can also represent this by electron dot structure,



v. Write the names of two types of bond?

Ans: Types of bonds:

Depending on the tendencies of atoms to lose or gain or share electrons, there are two types of bonds.

i. Ionic bonds   ii. Covalent bonds

vi. Write one main difference between an ionic and a covalent bond.

Ans: Difference between an ionic and a covalent bond:

Electrovalent (Ionic) Bond	Covalent Bond
1. It is formed by the complete transfer of electrons from one atom to the other.	It is formed by the mutual sharing of equal number of electrons between two atoms
2. It is indicated by positive and negative charges.	It is represented by a small straight line.
3. These are non-directional.	These are directional.
4. In ionic bond electronegativity difference between two atoms is > **1.7**	In Covalent bond electronegativity difference between two atoms is < **1.7**
Examples: $NaCl$, ,$CaCl_2$, K_2O, MgO	**Examples:** H_2 ,Cl_2, O_2, N_2

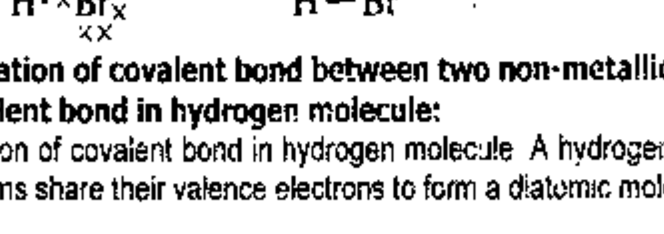
vii. What is a polar covalent bond? Give two examples of the compounds containing polar covalent bond?

Ans: Polar covalent bond:

The covalent bond formed between two unlike atoms which differ in their electronegativity is said to be a polar covalent bond.

Explanation:

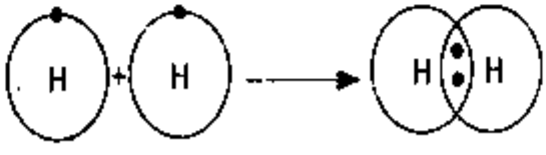
When two different atoms share electron pair, both the atoms exert different forces on the shared electron pair. More electronegative atom pulls shared electrons pairs with greater force then the other. So more electronegative atom partially draws electron density toward itself. This makes it partially negatively charged and other atom partially positively charged. Such a covalent bond is called polar covalent bond



viii. Describe the formation of covalent bond between two non-metallic elements.

Ans: Formation of covalent bond in hydrogen molecule:

Consider the formation of covalent bond in hydrogen molecule. A hydrogen atom has a single valence electron. Two hydrogen atoms share their valence electrons to form a diatomic molecule



In the formation of this molecule, each hydrogen atom achieves the electron configuration of the noble gas, helium which has two valence electrons

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. **(5 × 3 = 15)**

i. What is the total number of shared electrons in a molecule of CO_2?

Ans: $\ddot{O}::C::\ddot{O}$

Total number of shared electrons in $CO_2 = 4 + 4 = 8$

ii. Describe the factors affecting the formation of an ionic bond?

Ans: Factors Affecting the Formation of an Ionic Bond:

Following conditions favor the formation of an ionic bond.

(i) Low ionization energy:

Lower the ionization energy, more favourable is the formation of cation from an atom. Alkali metals have low values of ionization energy and hence they form ionic bonds.

(ii) High electron affinity:

Higher electron affinity should favor the formation of an anion. Halogens have high electron affinities and, therefore, they generally form ionic compounds.

(iii) Large lattice energy:

In ionic compounds, cations and anions attract each other due to the **coulombic force of attraction.** Due to these attractions, certain amount of energy is released. If the coulombic attractions are stronger then more energy is released and a more stable ionic bond is formed.

iii. What is meant by intermolecular force? Describe their types?

Ans: Intermolecular forces are the attractive forces between the molecules:

Types of Intermolecular forces:

The intermolecular forces are of four types and are collectively called Van der Waal's forces.

i. Dipole-dipole forces   ii. Ion-dipole forces
iii. Dispersion forces (London forces)   iv. Hydrogen bonding

iv. What is the total number of shared electrons in a molecule of CO_2?

Ans: $\ddot{O}::C::\ddot{O}$ Total number of shared electrons in $CO_2 = 4 + 4 = 8$

v. Discuss the nature and consequences of intermolecular forces?

Ans: Intermolecular forces:

The forces of attraction created between the molecules are called intermolecular forces.

For example:

$$H^{+\delta}-Cl^{-\delta}, H^{+\delta}-\underset{\underset{H^{\delta}}{|}}{O^{-\delta}} \quad \text{etc.}$$

Consequences of intermolecular forces:

Polar and non polar covalent intermolecular forces are weaker than an ionic or a covalent bond. There are several types of intermolecular forces

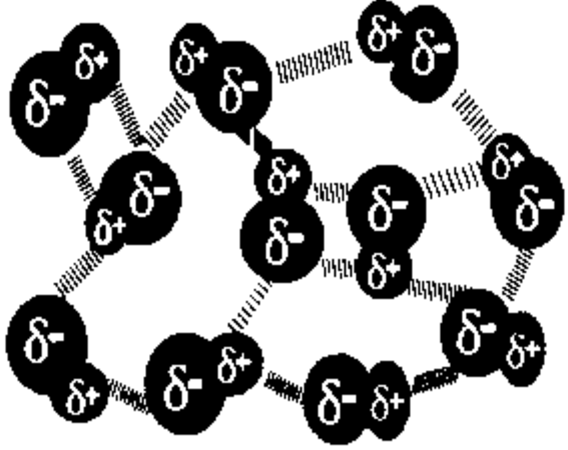
Paints and dyes are used to protect solid surfaces from the atmospheric effects. They also give visual appeal. Resins are used to coat materials that give toughness, flexibility, adhesion and chemical resistance For example dams, bridges, floors, trains, buses, cars etc are painted with resins. The synthetic resins are used where water resistance is required. Chemically, resins are either adhesive or they form bond linkages with the material being bonded together.

These intermolecular forces are extremely important in determining properties of water, biological molecules, such as proteins, DNA etc and synthetic materials such as glue, paints, resins etc. The adhesive

vi. Define dipole-dipole interactions. What is the nature of these linkages?

Ans: Dipole-Dipole interactions:

Slightly negative end of polar molecule is weakly attracted to the slightly positive end of another molecule. Such attracting forces are called dipole-dipole interactions. These linkage are weak and temporarily.



Dipole-Dipole interactions

vii. Describe the ways in which bonds may be formed.

Ans: The two main types of bonds formed between atoms are ionic bonds and covalent bonds.

An **ionic bond** is formed when one atom accepts or donates one or more of its valence electrons to another atom. A **covalent bond** is formed when atoms share valence electrons. The atoms do not always share the electrons equally, so a polar covalent bond may be the result.

When electrons are shared by two metallic atoms a metallic bond may be formed.

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. **(2 × 10 = 20)**

Q.4 a. How is an ionic bond formed? Explain with example.

Ans: Ionic Bonds:

Ionic bonds are formed between two atoms, when one atom loses electrons and other atom gains these electrons. The force of attraction that binds oppositely charged ions are called ionic bonds.

Explanation: Ionic bonds are formed between two atoms, when one atom loses electron to form cation and the other atom gains this electron to form anion.

Ionic bonds:

Anions and cations have opposite charges. They attract one another by electrostatic forces. **"The forces of attraction that bind oppositely charged ions are called ionic bonds"**. Compounds that consist of ions joined by electrostatic forces are called ionic compound. The total positive charges of the cations must be equal to the total negative charges of the anions. This is because ionic compounds as a whole are electrically neutral.

Examples of Formation of Ionic Bond:

(i) Formation of sodium chloride (NaCl):

The electronic configuration of sodium ($_{11}Na$) atom is 2, 8, 1. It has one electron in its valence shell. The electronic configuration of chlorine ($_{17}Cl$) atom is 2, 8,7. It has 7 valence electrons.

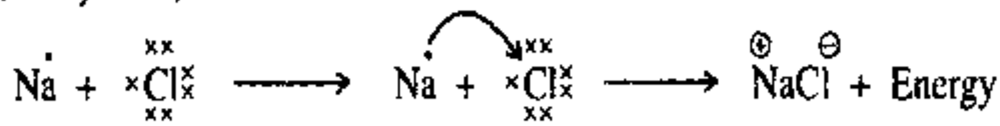
When the two atoms approach each other, there is a transfer of one electron from sodium (**$_{11}Na$**) to chlorine (**$_{17}Cl$**).

$$\underset{\substack{(2,\,8,1)\\ \text{Sodium atom}}}{Na} \xrightarrow{\text{Loss of electron } (e^-)} \underset{\substack{(2,\,8)\\ \text{Sodium ion}\\ \text{(Neon configuration)}}}{Na^+} + \text{electron } (e^-)$$

$$\underset{(2,\,8,\,7)}{Cl} + \text{electron } (e^-) \xrightarrow{\text{Gain of electron } (e^-)} \underset{(2,\,8,\,8)}{Cl^-}$$

In this way, both the atoms attain noble gas configuration The two ions, Na^+ and Cl^- are then held together by the coulombic force of attraction.

In terms of Lewis symbols, it can be written as follows:



Overall energy is released during the formation of NaCl.

Note that when atoms lose electrons, they become smaller and when the gain electrons they become larger.

Examples of ionic bond: **MgO, NaCl, $CaCl_2$, CaF_2, KCl, CaO**

Characteristics of ionic Bond:

(i) The atom which loses its electrons becomes positively charged ion, called **cation.** The atom which gains these electrons becomes negatively charged ion called **anion.**

(ii) These two oppositely charged ions are then held together by the **electrostatic forces of attraction** to form an ionic bond

(iii) Those compounds in which the constituent ions are held together by ionic bonds are called **ionic compounds.**

(iv) Ionic bond is produced between an element having low ionization energy and other having high electron affinity.

(v) The elements of group 1 and 2 form ionic compounds when react with elements of groups 16, 17.

b. What is a non polar covalent bond? Give two examples of the compounds containing non polar covalent bond?

Ans: Non polar covalent bond:

When two identical atoms share electron pairs, both the atoms exert same force on the shared electron pairs. Such a covalent bond is called non-polar covalent bond.

Examples: For example, bond in **H-H, Cl - Cl, O = O, N≡N** etc are non-polar covalent bonds.

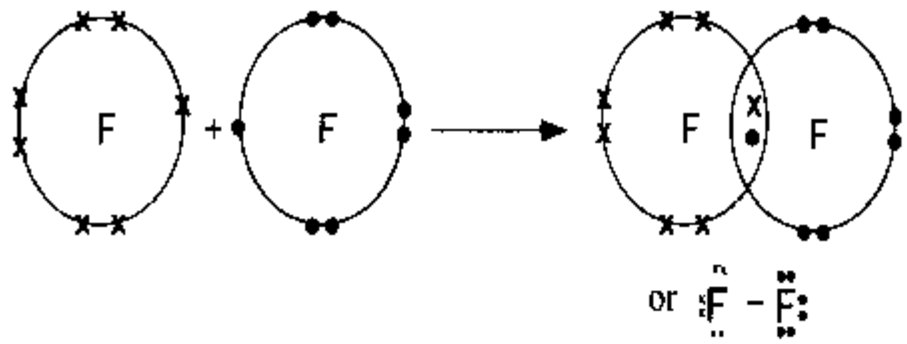
Q.5 a. What is covalent bond? Describe the various types of covalent bonds. Give one example of each kind?

Ans: Covalent Bond: A Covalent bond is formed by mutual sharing of electrons between two atoms.

Types of covalent bond:

i. Single covalent bond:

Covalent bond that is formed by the sharing of one electron pair is called single covalent bond. So H_2 and F_2 molecules contain a single covalent bond.

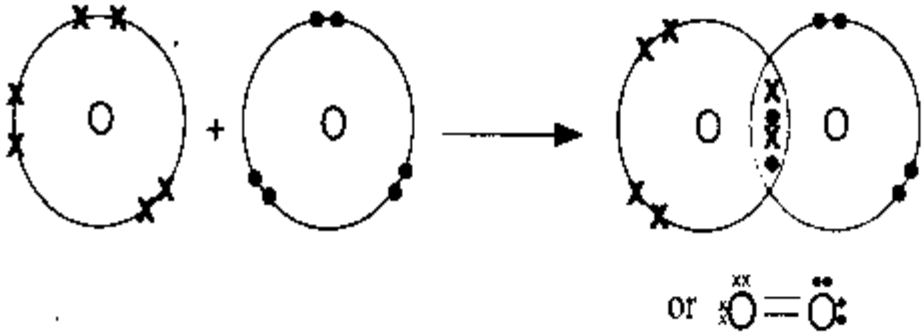


Lone pairs:

Pairs of valence electrons which are not shared between atoms are called unshared pairs or lone pairs.

Examples: **Cl−Cl, H−H, Br−Br, CH_3-CH_3 (Ethane)**

Double covalent bond:



Examples: O=O, $CH_2=CH_2$ (Ethene), O=C=O (Carbon dioxide)

Triple covalent bond: Triple covalent bonds are the bonds that involve three shared pairs of electrons.

Examples: N≡N, HC≡CH (Ethyne)

b. What is meant by hydrogen bonding? Discuss the nature and consequences of hydrogen bonding.

OR

Describe hydrogen bond and how hydrogen bonding is different from ionic and covalent bonding?

Ans: Hydrogen bonding:

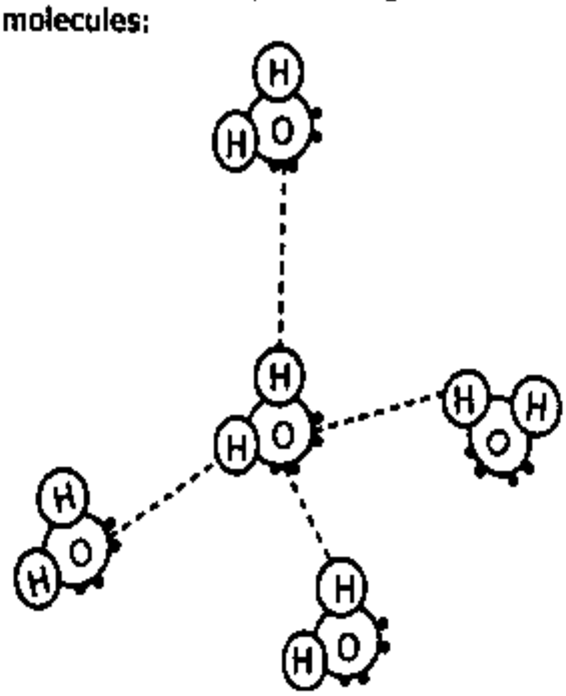
The interaction of a highly electron deficient hydrogen and lone pair on a nearby highly electronegative atom such as N, O or F is called hydrogen bond This phenomenon is called hydrogen bonding.

Explanation:

Hydrogen bonding is an intermolecular force. Molecules in which hydrogen is covalently bonded to a very electronegative atom such as oxygen, nitrogen or fluorine is also weakly bonded to a lone pair of electron of another electronegative atom. This other atom may occur in the same molecule or in a nearby molecule. This intermolecular interaction is called hydrogen bonding.

Oxygen, nitrogen or fluorine makes hydrogen very electron-deficient Thus interaction of such a highly electron deficient hydrogen and lone pair on a nearby electronegative atom compensates for the deficiency.

Hydrogen bonding in water molecules:

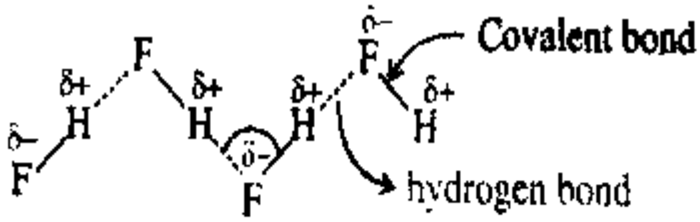


Hydrogen bonding in water

Examples of compounds showing hydrogen bonding:

(i) Hydrogen fluoride (HF):

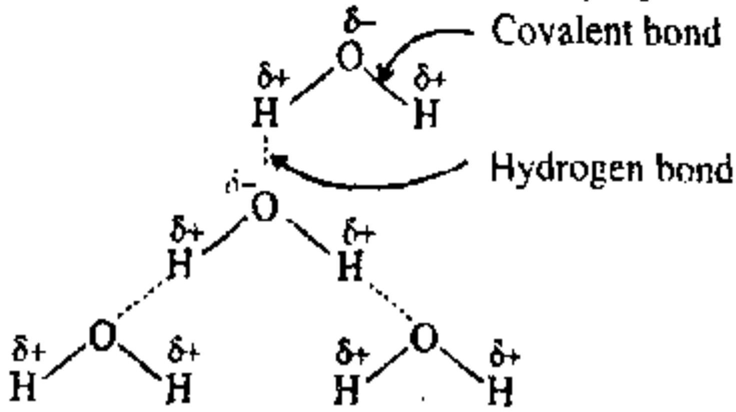
The H-bond in HF is so strong that the molecules remain associated through



Hydrogen bonding in HF

The average number of H-bonded molecules of hydrogen fluoride to produce a big molecule is six and so the formula is $(HF)_6$.

(ii) **Water (H_2O)**: Water is a typical polar compound which exhibits hydrogen bonding as shown below in



Hydrogen bonding in H_2O.

It is due to hydrogen bonding that water exists in the associated form.

Nature of hydrogen Bonding: Hydrogen bonding is intermolecular force.

Note: The adhesive action of paints and dyes is developed due to hydrogen bonding.

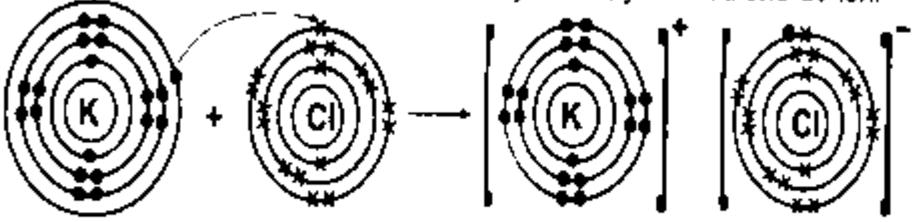
Q.6 a. For each of the following pairs of atoms, use electron dot and electron cross structures to write the equation for the formation of ionic compound.

(i) K and Cl **(ii) Ca and S** **(iii) Al and N**

Ans: (i) K and Cl:

K is metal and Cl is non-metal.

Metal atom tends to lose electrons and non – metal atoms tends to gain electrons to acquire electronic configuration of nearest noble gas Since K atom has one electron in the outer most shell. It losses one electron to form K^+ ion. Since Cl atom has seven electrons in outermost shell, it needs one electron to complete octet. So it gains one electron to form Cl^- ion. For every K^+ ion, you need one Cl^- ion.



(ii) Ca and S:

Ca is metal and S is non-metal.

Ca atom has two electrons in outermost shell. It losses two electrons to form Ca^{+2} ion. Since S atom has six electrons in outermost shell, so it gains two electrons to form S^{-2} ion.

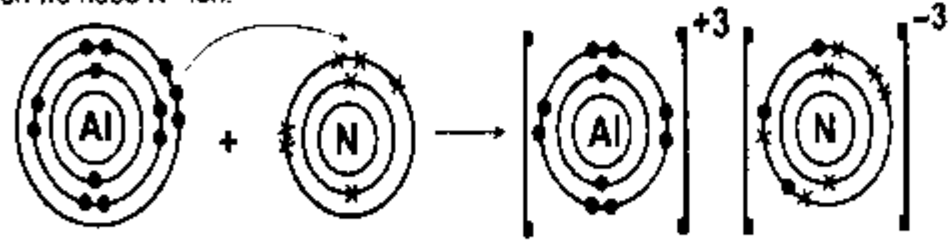
For every Ca^{+2} ion we need S^{-1} ions.

(iii) Al and N:

Al atom has three electrons in outermost shell. It losses three electrons to form Al^{+3} ion.

Since N atom has five electrons in outermost shell, so it gains three electrons to form N^{-3} ion.

For every Al^{+3} ion we need N^{-3} ion.



b. Recognize the following compounds as having ionic bonds.

(i) $MgCl_2$ (ii) KBr (iii) NaI

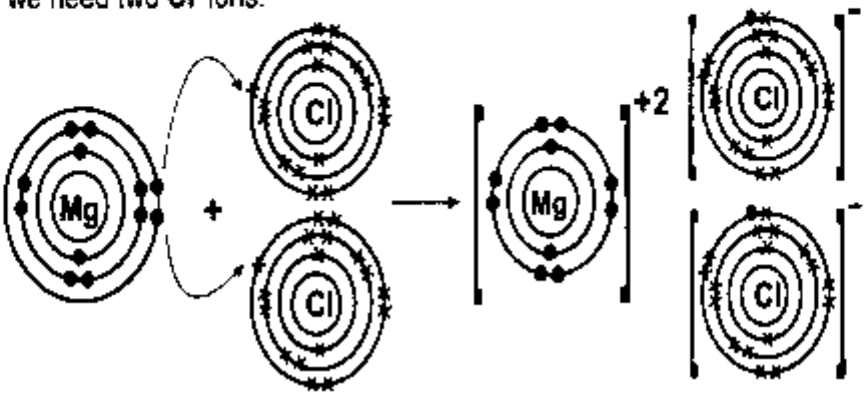
Ans: (i) $MgCl_2$:

Mg is metal and Cl is non-metal.

Mg atom has two electrons in outermost shell. It losses two electrons to form Mg^{+2} ion.

Since Cl atom has seven electrons in outermost shell, so it gains one electron to form Cl^- ion.

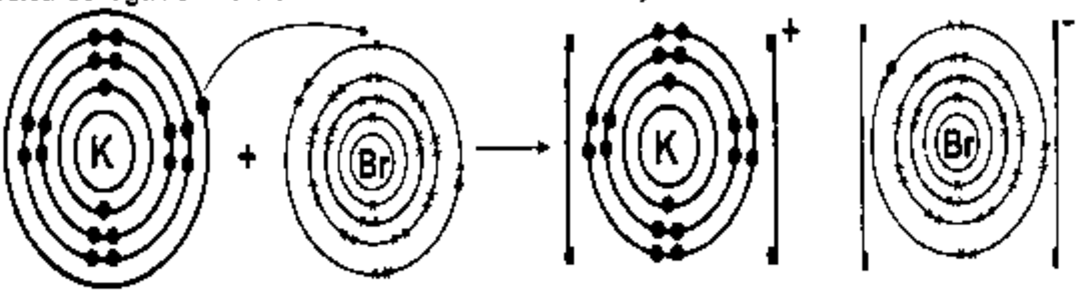
For every Mg^{+2} ion we need two Cl^- ions.



(ii) KBr:

K is metal and Br is non-metal.

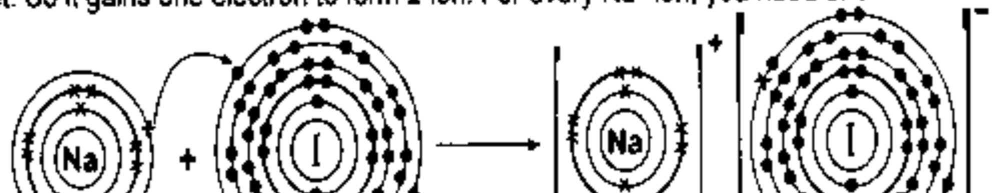
Metal atom tends to lose electrons and non – metal atoms tends to gain electrons to acquire electronic configuration of nearest noble gas. Since K atom has one electron in the outer most shell. It losses one electron to form K^+ ion. Since Br atom has seven electrons in outermost shell, it needs one electron to complete octet. So it gains one electron to form Br^- ion. For every K^+ ion, we need one Br^- ion.



(iii) NaI:

Na is metal and I is non-metal.

Metal atom tends to lose electrons and non – metal atoms tends to gain electrons to acquire electronic configuration of nearest noble gas. Since Na atom has one electron in the outer most shell. It losses one electron to form Na^+ ion. Since I atom has seven electrons in outermost shell, it needs one electron to complete octet. So it gains one electron to form I^- ion. For every Na^+ ion, you need one I^- ion.



IMPORTANT SHORT QUESTION ANSWERS
(Reduced Syllabus)

SELF ASSESSMENT EXERCISE 4.4

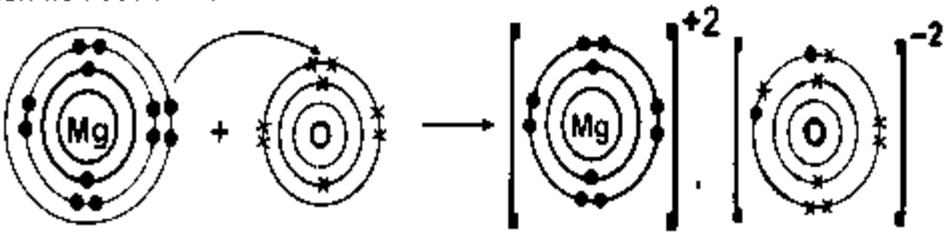
For each of the following pairs of atoms, use electron dot and electron cross structures to write the equation for the formation of ionic compound. **(a)** Mg and O **(b)** Al and Cl

Solution: **(a)** Mg and O

Mg atom has two electrons in outermost shell. It losses two electrons to form Mg^{+2} ion.

Since O atom has six electrons in outermost shell, so it gains two electrons to form O^{-2} ion.

For every Mg^{+2} ion we need two O^{-2} ions.



(b) Al and Cl

Al is metal and Cl is non-metal. Al atom has three electrons in outmost shell. So it loses three electrons to form Al^{+3} ion. Since Cl atom has seven electrons in outermost shell, so it gains one electron to form Cl^- ion. Al atom by losing three electrons and 3Cl atom by gaining one electron acquire nearest noble gas electronic configuration. We need three Cl^- ions for each Al^{+3} ion. Therefore, $AlCl_3$ is an ionic compound.

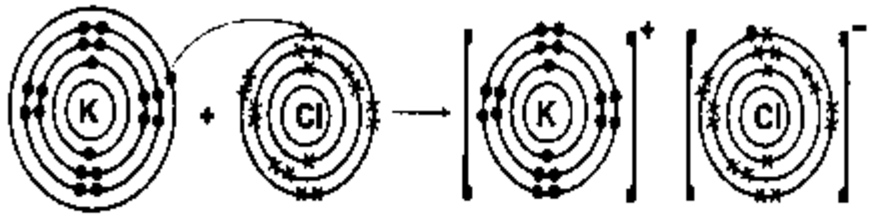
SELF ASSESSMENT EXERCISE 4.5

Recognize the following compounds as having ionic bonds:

(a) KCl **(b)** $AlCl_3$ **(c)** MgF_2

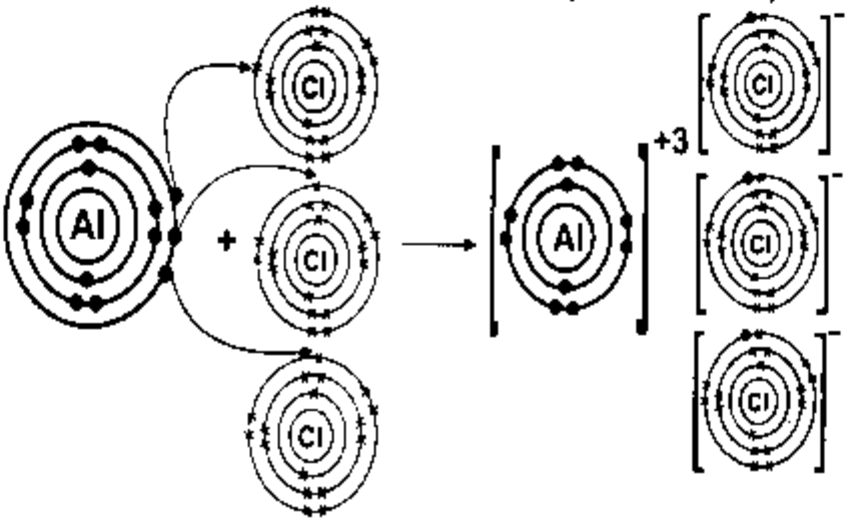
Solution: **(a)** KCl:

K is metal and Cl is non-metal. K atom has one electron in outmost shell. So it loses one electron to form K^+ ion. Since Cl atom has seven electrons in outermost shell, so it gains one electron to form Cl^- ion. K atom by losing one electron and Cl atom by gaining one electron acquire nearest noble gas electronic configuration. We need one Cl^- ion for each K^+ ion. Therefore, KCl is an ionic compound.



(b) $AlCl_3$:

Al is metal and Cl is non-metal. Al atom has three electrons in outmost shell. So it loses three electrons to form Al^{+3} ion. Since Cl atom has seven electrons in outermost shell, so it gains one electron to form Cl^- ion. Al atom by losing three electrons and 3Cl atom by gaining one electron acquire nearest noble gas electronic configuration. We need three Cl^- ions for each Al^{+3} ion. Therefore, $AlCl_3$ is an ionic compound.



(c) MgF_2:

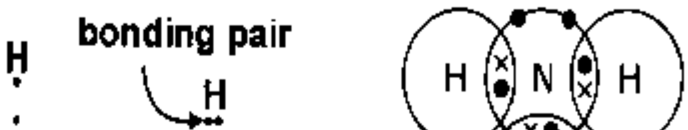
Mg is metal and F is non-metal. Mg atom has two electrons in outmost shell. So it loses two electrons to form Mg^{+2} ion. Since F atom has seven electrons in outermost shell, so it gains one electron to form F^- ion. Mg atom by losing two electrons and F atom by gaining one electron acquire nearest noble gas electronic configuration. We need two F^- ions for each Mg^{+2} ion. Therefore, MgF_2 is an ionic compound.

SELF ASSESSMENT EXERCISE 4.6

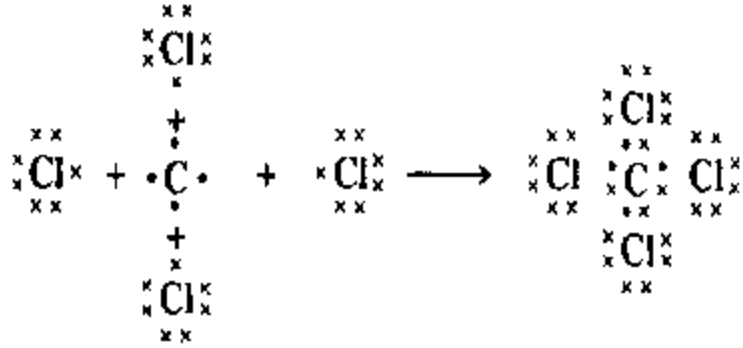
Draw electron cross and dot structures for the following molecules:

(a) NH_3 that is used to manufacture urea. **(b)** CCl_4, a dry cleansing agent.
(c) $SiCl_4$, used to make smoke screens. **(b)** H_2S, a poisonous gas.

Solution: **(a)** NH_3 that is used to manufacture urea.

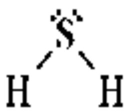


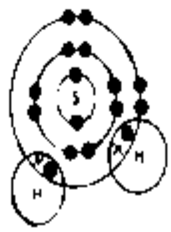
(b) CCl_4, a dry cleansing agent.



(c) $SiCl_4$, used to make smoke screens.

(b) H_2S, a poisonous gas.



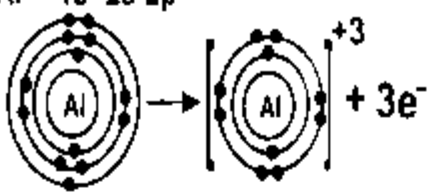


Q.2 Represent the formation of cations for the following metal atoms using electron dot structures. (a) Al (b) Sr (c) Ba

Ans: (a) Formation of Al^{+3} ion: (Atomic number of Aluminum = 13)

Since Al atom has three electrons in the outer most shell It losses three electrons to form Al^{+3} ion.

$$Al\ 1s^2\,2s^22p^63s^23p^1 \xrightarrow{-3e^-} Al^{+3}\ 1s^2\,2s^22p^6$$

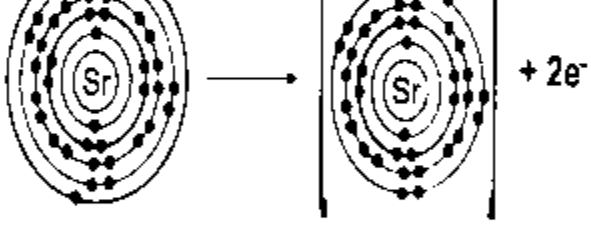


(b) Formation of Sr^{+2} ion: (Atomic number of strontium = 38)

Since Sr atom has two electrons in the outer most shell It losses two electrons to form Sr^{+2} ion.

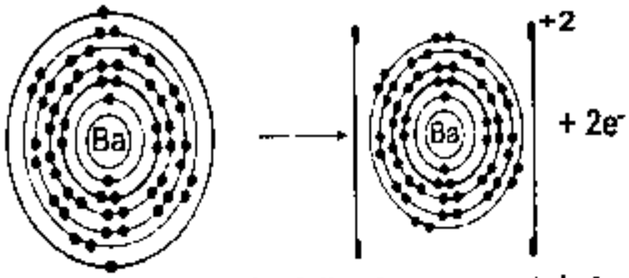
$$(2,8,18,8,2) \xrightarrow{-2e^-} (2,8,18,8)$$

$$Sr\ 1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2 \xrightarrow{-2e^-} Sr^{+2}\ 1s^2\,2s^22p^63s^23p^64s^23d^{10}4p^6$$



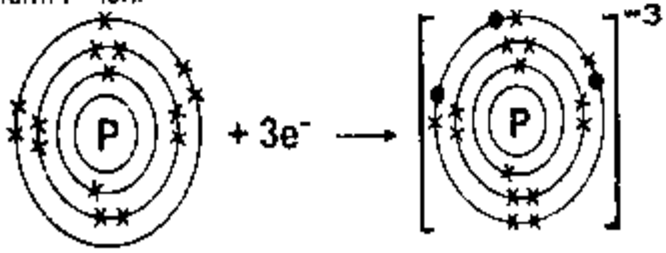
(c) Formation of Ba^{+2} ion: (Atomic number of barium = 56)

Since Ba atom has two electrons in the outer most shell It losses two electrons to form Ba^{+2} ion.

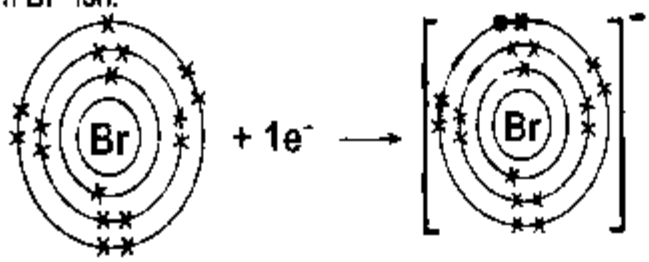


Q.3 Describe the formation of anions for the following non-metal atoms:
(a) P (b) Br (c) H

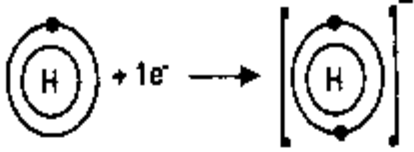
Ans: (a) Formation of P^{-3} anion:
Since P atom has five electrons in outermost shell, it needs three electrons to complete octet. So it gains three electrons to form P^{-3} ion.



(b) Formation of Br^{-1} anion:
Since Br atom has seven electrons in outermost shell, it needs one electron to complete octet. So it gains one electron to form Br^{-1} ion.



(c) Formation of H^{-1} anion:
Since H atom has one electron in outermost shell it needs one electron to complete duplet. So it gains one electron to form H^{-1} ion.

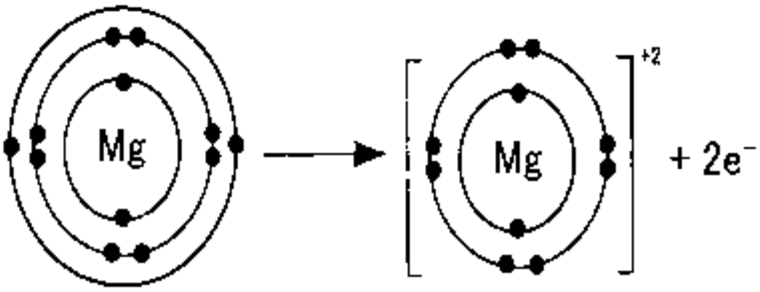


Q.4 Represent the formation of cations for the following metal atoms using electron dot structures. (a) Mg (b) Li (c) Be

Ans: (a) Formation of Mg^{+2} ion. (Atomic number of Magnesium = 12)

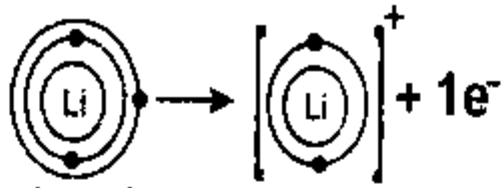
$$Mg\ 1s^2 2s^2 2p^6 3s^2 \xrightarrow{-2e^-} Mg^{+2}\ 1s^2 2s^2 2p^6$$

We can also represent this by electron dot structure,



(b) Formation of Li^+ ion. (Atomic number of Lithium = 3)

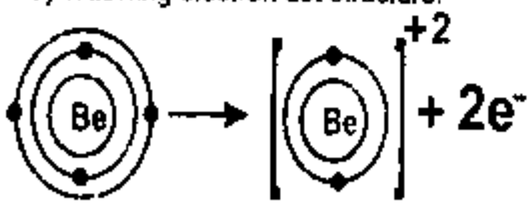
$$Li\ 1s^2 2s^1 \xrightarrow{-1e^-} Li^{+1}\ 1s^2$$



(c) Formation of Be^{+2} ion. (Atomic number of Beryllium = 4)

$Be \quad 1s^2 2s^2 \xrightarrow{-2e^-} Be^{+2} \quad 1s^2$

We can also represent Be^{+2} by following electron dot structure.



Q.5 An atom of an element has atomic number 9 and mass number 19.

(a) State the number of protons and neutrons in the nucleus of this atom.

(b) State the number of electrons in this atom.

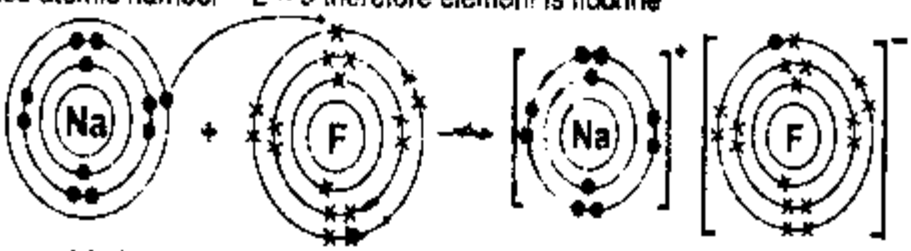
(c) Show with electron cross-dot diagrams, the formation of ions in the reaction of this atom with sodium atom.

Ans: **a.** Number of protons = Atomic number = Z = 9

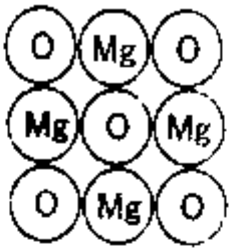
Number of protons = Atomic mass – Atomic number = A - Z = 19 – 9 = 10

b. Number of electrons = Atomic number = Z = 9

c. Since atomic number = Z = 9 therefore element is fluorine



Q.6 Magnesium oxide is a compound made up of magnesium ions and oxide ions.



(a) What is the charge on these ions.

(b) How these ions get these charges.

(c) Show with electron cross-dot diagrams the formation of these ions.

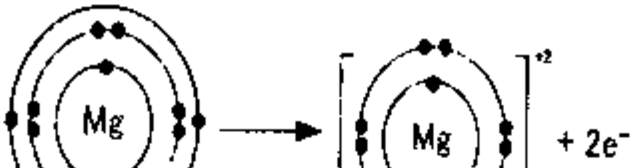
Ans: (a) Mg^{+2} and O^{-2}

(b) Mg^{+2} ion is formed by losing two electrons whereas O^{-2} ion is formed by gaining two electrons.

(c) Formation of Mg^{+2} ion.

$Mg \quad 1s^2 2s^2 2p^6 3s^2 \xrightarrow{-2e^-} Mg^{+2} \quad 1s^2 2s^2 2p^6$

We can also represent this by electron dot structure.

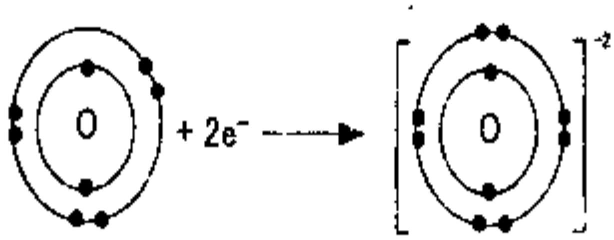


Formation of anion by oxygen atom (O^{-2}).

Oxygen belongs to Group VIA on the periodic table. So it has six electrons in its valence shell. It needs two electrons to achieve noble gas configuration.

$$\text{O} \quad 1s^2 2s^2 2p^4 + 2e^- \longrightarrow O^{-2} \;\; 1s^2 \underbrace{2s^2 2p^6}_{\text{octet}}$$

We can also represent this by electron dot structure.

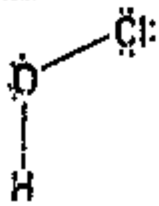


Q.7 Draw electron cross and dot structure for the following molecules:

(a) $COCl_2$, a poisonous gas called phosgene that has been used in World War-II.

(b) HOCl, hypochlorous acid is unstable, decomposes to liberate atomic oxygen that makes HOCl a strong oxidizing agent.

Ans:

(a) Phosgene $COCl_2$ (Carbonyl chloride):	**(b) HOCl (Hypochlorous acid):**
	**Bleach component:** 

Q.8 Electronic configuration of two elements X and Y are given below: $X = 1s^2 2s^2 2p^6 3s^2$
$Y = 1s^2 2s^2 2p^4$

Which of the following compounds is likely to form when X and Y react? Explain.

(a) A covalent compound of formula XY_2

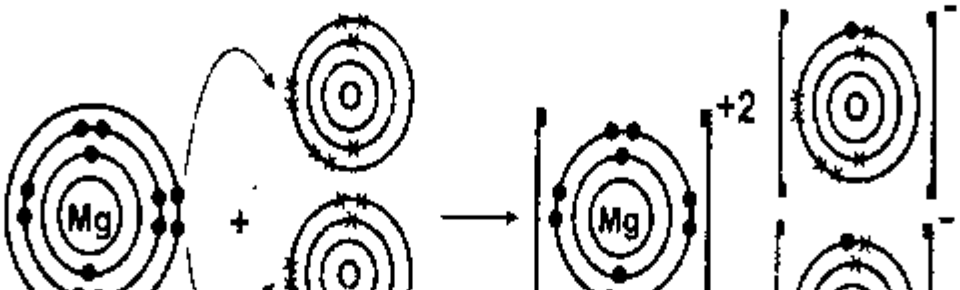
(b) An ionic compound of formula XY_2

(c) An ionic compound of formula XY

(d) An ionic compound of formula X_2Y

Solution: Option (C) is the correct answer. Because X loses 2 electrons from valence shell to form X^{+2}. Whereas Y gains 2 electrons to complete its octet and from Y^{-2}. So ionic compound of formula XY is formed.

$X = 1s^2 2s^2 2p^6 3s^2$ (Atomic number of element X = 12, the name of element is magnesium)
$Y = 1s^2 2s^2 2p^4$ (Atomic number of element Y = 8, the name of element is oxygen)



GUESS PAPER & MODEL PAPER # 5
BASED ON CHAPTER # 5 (Reduced Syllabus)
PHYSICAL STATES OF MATTER

CHAPTER 5: PHYSICAL STATE OF MATTER

Boyle's Law, Charles Law, Liquid state, evaporation, vapor pressure, boiling point, effect of external pressure on boiling point, sublimation, types of solids (Amorphous and crystalline).

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) Which statement about the particles of a solid is not correct?
A. They move at great speeds. B They are arranged in regular patterns.
C. There is a very little space between the particles
D. The forces of attraction between the particles are strong.

(ii) Conversion of a liquid to a gas at all temperatures is called
A. sublimation B. evaporation C. condensation D. boiling

(iii) At Mount Everest, the water boils at
A. 100 °C B. 90 °C C. 80 °C D. 70 °C

(iv) In which of the following processes speed of the particles decreases?
A. melting B. boiling C. sublimation D. condensation

(v) The mathematical expression that describes Boyle's law is
A. PV = constant B V × constant = P
C. P × constant = V D. V/P = constant

(vi) A liquid boils when its vapour pressure becomes equal to ___________.
A. 760 cm Hg B. 1 Pa C. 101.325 kPa D. 0.1 atm.

(vii) Acetone has characteristic fragrant odour and is used to make finger nail polish. It melts at -94 °C and boils at 56 °C. What is the physical state of acetone at 25 °C and 1atm?
A gas B. liquid C. solid D. cannot be predicated

(viii) Water normally boils at 100 °C, but it is possible for water to boil at room temperature. What variable would you have to change to do this.
A. increase external pressure B decrease external pressure
C. decrease temperature D. none of these

(ix) Bromine has a melting point of -7 °C and a boiling point of 59 °C. What is the physical state of bromine at 100 °C.
A. gas B. liquid C. solid D. cannot be predicted

(x) Which is not the property of crystalline solids?
A. have well defined shape B. have orderly arrangements of particles

(xi) 1atm = ________ kPa
A. 101.325 kPa B. 101.324 kPa C. 101.322 kPa D. 101.323 kPa

(xii) The vapour pressure at water at 20 °C is ________ kPa.
A. 1.32 B. 2.33 C. 2.53 D. 7.37

Time allowed: 2:40 hours Total Marks: 53

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet–B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. Name the factors that affect the rate of evaporation.
ii. How does temperature effect vapour pressure of a liquid?
iii. Water boils at 120 °C in a pressure cooker, why?
iv. Is evaporation a cooling process?
v. Does evaporation take place at all temperatures?
vi. Express the pressure 400 mm Hg in kPa?
vii. Describe the effect of temperature on vapour pressure?
viii. Does vapour pressure depend upon the nature of a liquid?

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. Give reason:
   a. When you put nail polish remover on your palm, you feel a sensation of coldness.
   b. Wet clothes dry quickly in summer than in winter.
ii. Explain the term boiling point.
iii. a. Does boiling point depend upon external pressure?
   b. Why water boils at 70 °C on the top of Mount Everest and at 120 °C in a pressure cooker?
iv. How do you compare boiling and evaporation?
v. Explain the effect of external pressure on boiling point.
vi. In automobile engine the gaseous fuel-air mixture enters the cylinder and is compressed by a moving piston before it is ignited. If the initial cylinder volume is 990 cm^3. After the piston moves up, the volume is 90 cm^3. The fuel-air mixture initially has a pressure of 1.0 atm. and final pressure 11.0 atm. Do you think this change occurs according to the Boyle's law.
vii. Why does evaporation lower the temperature of a liquid?

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 a. State and explain Boyle's law.
   b. Explain the terms.
   (a) evaporation (b) vapour pressure (c) boiling point

Q.5 a. Explain the effect of external pressure on boiling point.
   b. Differentiate between amorphous and crystalline solids.

SOLUTION OF GUESS PAPER & MODEL PAPER # 5
(Reduced Syllabus)

SECTION- A (MCQs)

i. A	ii. B	iii. D	iv. D	v. A	vi. C
vii. B	viii. B	ix. A	x. D	xi. A	xii. B

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. Name the factors that affect the rate of evaporation.

Ans: i. **Strength of intermolecular forces:**
Stronger the intermolecular force lower will be the rate of evaporation and vice versa.
ii. **Temperature:** Increase in temperature increases the rate of evaporation and vice versa.
iii. **Surface area:** Larger the surface area, higher is the rate of evaporation and vice versa.

ii. How does temperature effect vapour pressure of a liquid?

Ans: The vapour pressure of a liquid increases as the temperature increases.
The molecules in the liquid are more energetic at higher temperatures, and more molecules can escape from the liquid phase into the gas phase.

iii. Water boils at 120 °C in a pressure cooker, why?

Ans: Pressure cooker is equipped with a valve that controls the pressure inside the pot.
This valve generally exerts a pressure of 2 atm. Therefore, the valve does not allow water vapours to escape until the pressure inside the pot reaches 2 atm. Because vapour pressure of water becomes 2 atm when the temperature reaches 120 °C. So water boils at 120 °C in a pressure cooker.

iv. Is evaporation a cooling process?

Ans: Evaporation is the cooling process:
According to kinetic theory, the temperature is a measure of average kinetic energy of the molecules of a liquid. During evaporation, the escape of high energy molecules from the surface of a liquid, lowers the average kinetic energy of the remaining molecules and therefore the temperature of the liquids falls down. Thus evaporation is a cooling process.

v. Does evaporation take place at all temperatures?

Ans: The molecules whose kinetic energies are greater than the average kinetic energy of the molecules, escape from the surface of the liquid. If temperature of the liquid is increased, rate of evaporation also increases. Anyhow evaporation take place at all temperatures and only the rate differs.

vi. Express the pressure 400 mm Hg in kPa?

Ans: 1atm = 760 mm = 101.325 kPa = 1.01325×10^5 Pa. 760 mm = 1.01325 kPa

$$1 \text{ mm} = \frac{1.01325}{760}$$

$$400 \text{ mm} = \frac{1.01325}{760} \times 400 = 0.00133 \times 400 = 0.533 \text{ kPa}$$

vii. Describe the effect of temperature on vapour pressure?

Ans: Effect of temperature on vapour pressure:
Table shows vapour pressures of some liquids at various temperatures.

Table: Vapour pressures of some liquids at various temperatures

Vapour Pressure (kPa) of Several Substances at Various Temperatures						
	0°C	20°C	40°C	60°C	80°C	100°C

We will find that the vapour pressure of liquids changes with temperature. This is because an increase in temperature of a liquid increases the kinetic energy of the molecules. As a result more of the molecules will have minimum kinetic energy needed to escape the surface of the liquid

viii. Does vapour pressure depend upon the nature of a liquid?

Ans: Yes, vapour pressure depend upon the nature of a liquid The vapour pressure of a liquid decreases when inter molecular forces increases.

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. Give reason:

a. When you put nail polish remover on your palm, you feel a sensation of coldness.

b. Wet clothes dry quickly in summer than in winter.

a. When you put nail polish remover on your palm, you feel a sensation of coldness.

Ans: The chemical nail polish remover (usually acetone) is very volatile. it evaporates very quickly To change from a liquid state, to a gaseous state takes energy. The warmth of your palm supplies the energy. The flow of heat is from your palm and into the liquid This removes heat from your palm and has the sensation of being cool.

b. Wet clothes dry quickly in summer than in winter.

Ans: At higher temperature, more molecules of a liquid are moving with high velocities. Thus more molecules escape from its surface. Thus evaporation is faster than at low temperature. That is why wet cloth dry quickly in the winter.

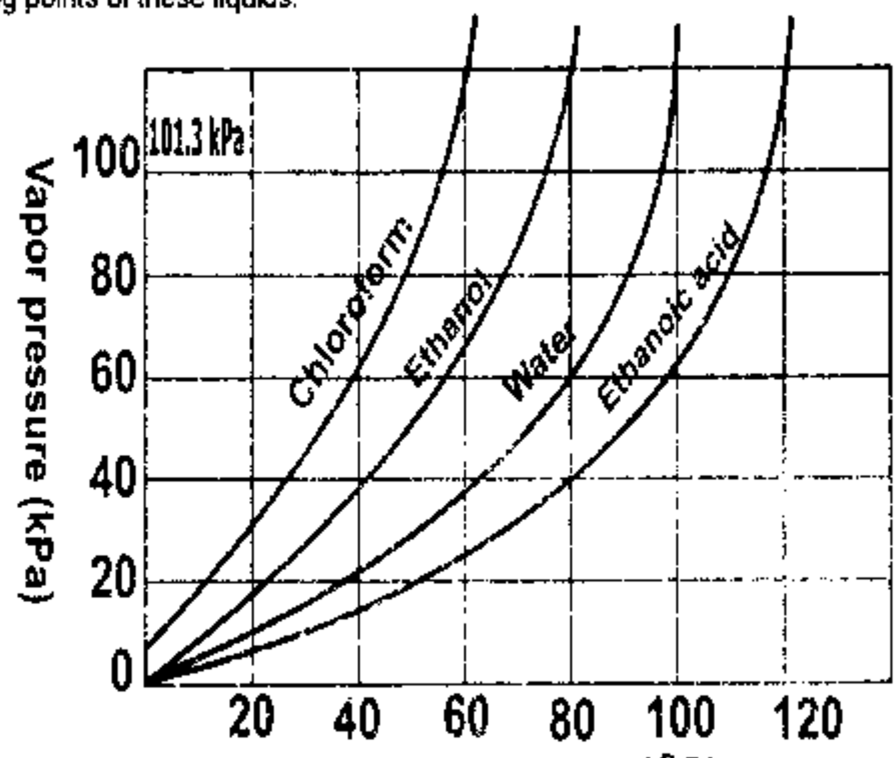
ii. Explain the term boiling point.

Ans: Boiling point:

The temperature at which vapour pressure of a liquid becomes equal to the external or atmospheric pressure is called boiling point.

Explanation:

Vapour pressure of the liquid keeps on increasing with the increase in temperature. At a certain temperature, the vapour pressure of the liquid becomes equal to the atmospheric pressure or external pressure. At this stage liquid starts boiling. Bubbles of vapours form throughout the liquid, rise to surface and escape into the air. Figure shows the variation in vapour pressures of four liquids with temperature. It also indicates the boiling points of these liquids.



iii. a. Does boiling point depend upon external pressure?

Ans: Yes, boiling point depends upon external pressure. The boiling point of a liquid decreases with the decrease in external pressure.

b. Why water boils at 70 °C on the top of Mount Everest and at 120 °C in a pressure cooker?

Ans: When the pressure of atmosphere is 1atm or 101.325 kPa water boils at 100 °C at sea level. This is because at this temperature vapour pressure of water is 1 atm or 101.325 kPa. At Mount Everest at about 8850m above sea level atmospheric pressure is only 34kPa So water boils at this height above sea level, when its vapour pressure is 34kPa at 70 °C. So water boils at 70 °C.

Pressure cooker is equipped with a valve that controls the pressure inside the pot This valve generally exerts a pressure of 2 atm. Therefore, the valve does not allow water vapours to escape until the pressure inside the pot reaches 2 atm. Because vapour pressure of water becomes 2 atm when the temperature reaches 120 °C. So water boils at 120 °C in a pressure cooker

iv.* How do you compare boiling and evaporation?

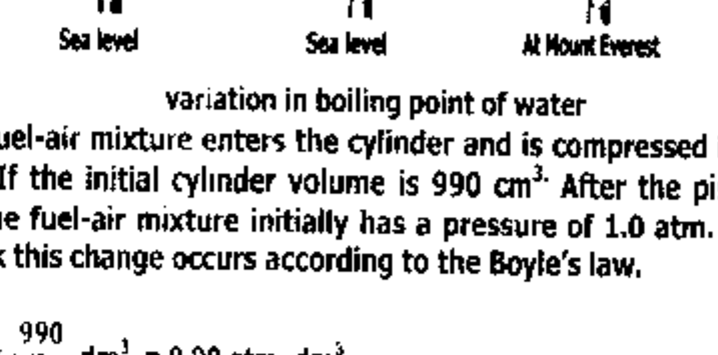
Ans: Comparison between boiling and evaporation:

Boiling	Evaporation
i. It takes place at a fixed temperature e.g. boiling of water is 100 °C.	i. it happens at any temperature.
ii. It happens throughout the bulk of liquid.	ii. It only takes place at the surface of the liquid.
iii. Bubbles are formed.	iii. No bubbles are produced.
iv. It is a quick process.	iv. It is a slow process.
v. It is affected at the altitude.	v. It is not affected by altitude.

v. Explain the effect of external pressure on boiling point.

Ans: Effect of external pressure on boiling point:

Liquids boil when their vapour pressure is equal to the pressure exerted on the liquid by its surroundings. The normal boiling point of water is 100°C. In the mountains the atmospheric pressure is less than 1 atm. so water boils below 100°C. In a pressure cooker at 2 atm, water does not boil until the temperature reaches 120°C.



variation in boiling point of water

vi. In automobile engine the gaseous fuel-air mixture enters the cylinder and is compressed by a moving piston before it is ignited. If the initial cylinder volume is 990 cm³. After the piston moves up, the volume is 90 cm³. The fuel-air mixture initially has a pressure of 1.0 atm. and final pressure 11.0 atm. Do you think this change occurs according to the Boyle's law.

Solution: ($1\ dm^3 = 1000\ cm^3$)

$$P_1 \times V_1 \text{ before change} = 1 \times atm \times \frac{990}{1000} dm^3 = 0.99\ atm.\ dm^3$$

$$P_2 \times V_2 \text{ after change} = 11\ atm \times \frac{90}{1000} dm^3 = 0.99\ atm.\ dm^3$$

$$P_1V_1 = P_2V_2$$

Thus the calculated result agrees with the pressure-volume relationship according to the Boyle's Law.

vii. Why does evaporation lower the temperature of a liquid?

Ans: Evaporation is the cooling process:

According to kinetic theory, the temperature is a measure of average kinetic energy of the molecules of

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. **(2 × 10 = 20)**

Q.4 a. State and explain Boyle's law.

Ans: Boyle's law: Boyle's law states that the volume of a fixed amount of a gas at a given temperature is inversely proportional to the applied pressure.

Mathematical form of Boyle's law:

$$p \propto \frac{1}{v}$$

$$p = \frac{1}{v} \times constant$$

$$V \times P = constant$$

This means product of volume and pressure must be constant if inverse relationship exists between them.

Experimental verification of Boyle's law:

Relationship between volume and pressure at constant temperature

Experiment No.	Pressure (atm)	Volume (dm^3)	V×P (dm^3.atm)
1.	0.500	4.00	2.00
2.	1.00	2.00	2.00
3.	2.00	1.00	2.00
4.	4.00	0.500	2.00

Table clearly shows that the product of volume and pressure is constant

Therefore, we can write $V \times P = \text{constant} \Rightarrow p = \frac{1}{v} \times \text{constant} \Rightarrow p \propto \frac{1}{v}$

This means the volume of a given mass of a gas at constant temperature is inversely proportional to the pressure on the gas. This relationship is known as Boyle's Law.

b. Explain the terms. (a) evaporation (b) vapour pressure (c) boiling point

Ans: a. Evaporation:

Conversion of a liquid to a gas or vapour at all temperatures is called vaporization or evaporation.

Explanation:

Place some liquid such as ether or acetone in an open container and observe. You will notice that the volume of the liquid gradually decreases and finally no more of the liquid is left. This is because liquids constantly change into gas or vapours even when the temperature is less than the boiling point of a liquid.

In evaporation, some molecules in the liquid break away and enter the gas or vapour state.

Evaporation is the cooling process:

Only those molecules which have greater kinetic energy than average can break away from the surface. This means the molecules with the highest kinetic energy escape first. The molecules in the liquid have a lower average kinetic energy than the molecules that have escaped. Therefore liquid's temperature will decrease. Therefore evaporation is a cooling process.

b. Vapour pressure:

The pressure exerted by the vapours of a liquid in equilibrium with its liquid is called vapour pressure.

Explanation:

In a closed container no molecules can escape into the outside air, when a partially filled container is sealed. Some of the liquid molecules vapourize. As the time passes, the number of molecules changing into vapours increases. Some of these molecules because of their random motions will collide with the liquid surface. Such molecules are recaptured by the molecules at the surface of the liquid. This process is called condensation. These two opposing processes will continue. After some time, the number of

$$\text{Liquid} \underset{\text{condensation}}{\overset{\text{evaporation}}{\rightleftharpoons}} \text{vapours}$$

At equilibrium liquid continues to evaporate and condense but at equal rates. At equilibrium the space above the liquid is saturated with vapours

c. Boiling point:

The temperature at which vapour pressure of a liquid becomes equal to the external or atmospheric pressure is called boiling point

Explanation:

Vapour pressure of the liquid keeps on increasing with the increase in temperature. At a certain temperature, the vapour pressure of the liquid becomes equal to the atmospheric pressure or external pressure. At this stage liquid starts boiling. Bubbles of vapours form throughout the liquid, rise to surface and escape into the air.

Note: Remember that a liquid boils when its vapour pressure becomes equal to the atmospheric pressure i.e., 760mmHg or 101.325 kPa at sea level.

Q.5 a. Explain the effect of external pressure on boiling point.

Ans: Effect of external pressure on boiling point:

Liquids boil when their vapour pressure is equal to the pressure exerted on the liquid by its surroundings. The normal boiling point of water is **100°C**. In the mountains the atmospheric pressure is less than 1 atm. so water boils below 100°C. In a pressure cooker at 2 atm, water does not boil until the temperature reaches 120°C

When the pressure of atmosphere is 1atm or 101.325 kPa water boils at **100 °C** at sea level. This is because at this temperature vapour pressure of water is 1 atm or 101.325 kPa At Mount Everest at about 8850m above sea level atmospheric pressure is only 34kPa. So water boils at this height above sea level, when its vapour pressure is 34kPa at 70 °C. So water boils at 70 °C

Pressure cooker is equipped with a valve that controls the pressure inside the pot. This valve generally exerts a pressure of 2 atm. Therefore, the valve does not allow water vapours to escape until the pressure inside the pot reaches 2 atm Because vapour pressure of water becomes 2 atm when the temperature reaches 120 °C. So water boils at **120 °C** in a pressure cooker

b. Differentiate between amorphous and crystalline solids.

Ans: Crystalline Solids:

Property		Crystalline solids
(i)	Definition	The solids in which atoms, molecules or ions are arranged in a regular repeating three-dimensional well-ordered pattern are called as crystalline solids.
(ii)	Geometry	These solids show characteristic geometrical shapes.
(iii)	Melting	These solids melt sharply at their melting points. Crystalline solids have sharp melting points.
(iv)	Order	These solids have perfect order
(v)	Examples	Sodium chloride (**$NaCl$**), silver nitrate (**$AgNO_3$**) are best examples of crystalline solids.

Amorphous Solids:

Property		Amorphous solids
(i)	Definition	Amorphous solids are those in which atoms ions or molecules are not arranged in a definite pattern, rather these are randomly arranged.
(ii)	Geometry	These solids generally appear in lump or fine powder form.
(iii)	Melting	Amorphous solids do not melt. They simply soften on heating, and gradually begin to flow on further heating. These solids are, therefore, considered as super cooled liquids.
(iv)	Order	In amorphous solids, perfect order is absent.

Q.6 a. Briefly explain the Charles's law.

Ans: Charles's law:

The law states that the volume of given mass of a gas varies directly with absolute temperature at constant pressure.

Mathematical form of Charles's law: $V \propto T$

$$V = \text{constant} \times T$$

$$\frac{V}{T} = \text{constant}$$

This relationship is known as the Charles's Law.

Table: Volume-Temperature relationship

Sr. No.	Volume (dm^3)	Temperature(K)	$\frac{V}{T}$
1.	1.00	273	0.0037
2.	1.37	373	0.0037
3.	1.73	473	0.0037
4.	2.10	573	0.0037

Since ratio of $\frac{V}{T}$ is constant, $\frac{V}{T}$ = constant

$V = T \times \text{constant}$ and $V \propto T$

b. Explain the process of sublimation?

Ans: Sublimation:

Some solids on heating change to vapours without passing through the liquid state. This process is called sublimation.

Examples: Substances like iodine, benzonic acid, ammonium chloride etc. sublime. Figure shows sublimation of violet-black crystals of iodine.

IMPORTANT SHORT QUESTION ANSWERS (Reduced Syllabus)

SELF ASSESSMENT EXERCISE 5.2

1. **A bacterial culture isolated from sewage produces 36.4 cm^3 of methane (CH_4) gas at 27 °C and 760mm Hg. This gas occupies 33.124 cm^3 at 0 °C and same pressure. Explain volume-temperature relationship from this data.**

Solution: Problem Solving Strategy:

1. According to the Charles's law, ratio of volume to absolute temperature is constant for any set of conditions.
2. Convert °C temperature to Kelvin temperature by adding 273.

Table: Temperature-volume relationship

Temperature (°C)	Volume (cm^3)	Temperature (K)	$\frac{V}{T}$
27	36.4	300	$\frac{36.4}{300} = 0.1$
0	33.124	273	$\frac{33.124}{273} = 0.1$

The ratio $\frac{V}{T} = 0.1$ is fairly constant. Thus volume of the gas varies directly with the absolute temperature as stated by the Charles's law.

2. **A perfect elastic balloon filled with helium gas has a volume of $1.25\times10^3\,dm^3$ at 1.00 atm and 25 °C on ascending to a certain altitude where temperature is 15 °C the volume of balloon becomes $1.208\times10^3\,dm^3$. Show that this data satisfies the Charles's law.**

Solution: Problem Solving Strategy:

1. According to the Charles's law ratio of volume to absolute temperature is constant for any set of conditions.
2. Convert °C temperature to Kelvin temperature by adding 273.
3. Find $\frac{V}{T}$ for each set of conditions and compare.

Table: Temperature-volume relationship

Temperature (°C)	Volume (dm^3)	Temperature (K)	$\frac{V}{T}$
25	1.25×10^3	298	$\frac{1.25\times10^3}{298} = 4.19$
15	1.208×10^3	288	$\frac{1.208\times10^3}{288} = 4.19$

The ratio $\frac{V}{T} = 4.19$ is fairly constant. Thus volume of the gas varies directly with the absolute temperature as stated by the Charles's law.

SELF ASSESSMENT EXERCISE 5.5

1. **The boiling point of water on the top of Mount Everest is 70 °C, while at Murree 98 °C. Explain this difference.**

Ans: When the pressure of atmosphere is 1atm or 101.325 kPa water boils at 100 °C at sea level. This is because at this temperature vapour pressure of water is 1 atm or 101.325 kPa. At Mount Everest at about 8850 m above sea level atmospheric pressure is only 34 kPa So water boils at this height above sea level when its vapour pressure is 34 kPa at 70 °C. So water boils at 70 °C.

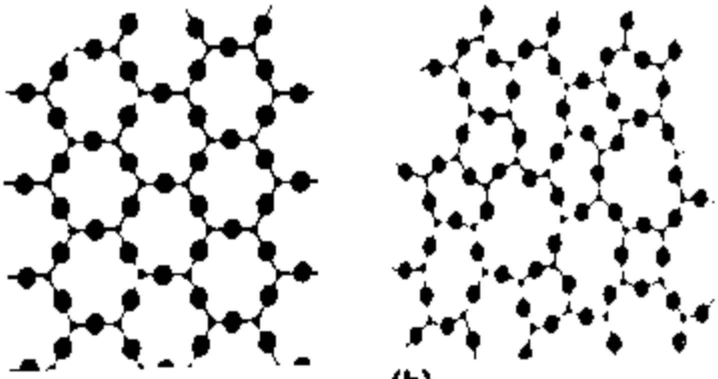
2. **If you try to cook an egg in boiling water while camping at an elevation of 0.5 km in the mountain, you will find that it takes longer than it does at home. Explain why?**

Ans: When the pressure of atmosphere is 1atm or 101.325 kPa water boils at 100 °C at sea level. This is because at this temperature vapour pressure of water is 1 atm or 101.325 kPa. At elevation of 0.5 km on the mountain above sea level atmospheric pressure becomes low So water boils at this height above sea level when its vapour pressure is low at low temperature. That is why egg takes longer time to cook than it does at home.

SELF ASSESSMENT EXERCISE 5.7

Quartz is the crystalline form of silicon dioxide (SiO_2). It is a hard, brittle and colourless solid. When quartz is heated above its melting point (about 1600°C) and cooled rapidly, an amorphous solid called quartz glass results.

Figure shows two-dimensional representation of quartz glass and quartz crystal. Identify each?



Two-dimensional representation of quartz glass and quartz crystal

Ans: "Figure (a)" represent quartz crystal because atoms are arranged in a regular repeating three-dimensional well-ordered pattern.
"Figure (b)" represent quartz glass because atoms are not arranged in a definite pattern, rather these are randomly arranged.

Q.1 The air in a perfectly elastic balloon occupies 885 cm^3, during the fall when the temperature is 20 °C. During the winter, the temperature on a particular day is -10 °C, the balloon occupies 794.39 cm^3. If the pressure remains constant. Show that the given data proves the volume temperature relation according to the Charles's Law.

Solution: Problem Solving Strategy:

1. According to the Charles's law, ratio of volume to absolute temperature is constant for any set of conditions
2. Convert °C temperature to Kelvin temperature by adding 273.
3. Find $\frac{V}{T}$ for each set of conditions and compare.

Table: Temperature-volume relationship

Temperature (°C)	Volume (cm^3)	Temperature (K)	$\frac{V}{T}$
20	885	293	$\frac{885}{293} = 3.0204$
-10	794 39	263	$\frac{794.39}{263} = 3.0204$

The ratio $\frac{V}{T} = 3.0204$ is fairly constant. Thus volume of the gas varies directly with the absolute temperature as stated by the Charles's law.

Q.2 A sample of neon that is used in a neon sign has a volume of 1500 cm^3 at a pressure of 636 torr. The volume of the gas after it is pumped into the glass tube of the sign is 1213.74 cm^3, when it shows a pressure of 786 torr. Show that this data obeys Boyle's law.

Solution: Problem Solving Strategy:

1. According to the Boyle's law, product of pressure and volume is constant at any two sets of conditions.
2. Calculate P × V for the two sets of condition and compare.

($1\ dm^3 = 1000\ cm^3$) , (1 atm = 760 torr)

$P_1 \times V_1$ before change $= \frac{636}{760} \times atm \times \frac{1500}{1000} dm^3 = 0.8368 \times atm \times 1.5\ dm^3$

$= 1.255\ atm.\ dm^3$

$P_2 \times V_2$ after change $= \frac{786}{760} \times atm \times \frac{1213.74}{1000} dm^3 = 1.0342 \times atm \times 1.21374\ dm^3$

$= 1.255\ atm.\ dm^3$ $\therefore$ $P_1V_1 = P_2V_2$

Thus the calculated result agrees with the pressure-volume relationship according to the Boyle's Law.

Q3. Identify the process occurring in each of the following:

a. Mothballs slowly disappear.

GUESS PAPER & MODEL PAPER # 6
BASED ON CHAPTER # 6 (Reduced Syllabus)
SOLUTIONS

CHAPTER 6 : PHYSICAL STATE OF MATTER

Saturated, Unsaturated, Super Saturated Solution, Concentration Units (Percentage, Molarity, Solubility, Effect of Temperature on Solubility, Comparison of Solution, Suspension and Colloids (Only table is included).

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) The maximum amount of sodium acetate that dissolves in 100 g of water at 0 °C is 119 g and 170 g at 100 °C. If you place 170 g of sodium acetate in 100 g of water at 0 °C, the resulting solution would be.
A. unsaturated B. saturated C. Supersaturated D. 1M

(ii) How many moles of sodium atoms are present in 2.3 g Na?
A. 1 B. 1.5 C. 0.1 D. 0.15

(iii) What is the mass of 5 moles of hydrogen gas?
A. 5 g B. 5.04 g C. 10.08 g D. 1.008 g

(iv) Solubility is the amount of solute dissolved in
A. 50 g of solvent B. 100 g of solvent
B. 250 g of solvent C. 200 g of solvent

(v) At a particular temperature, the solution which cannot dissolve more solute is called
A. saturated solution B. unsaturated solution
C. aqueous solution D supersaturated solution

(vi) A solution of NaOH has a concentration of 4 g/dm^3. What is the mass of NaOH contained in 250 cm^3 of this solution?
A. 40 g B 20 g C. 1 g D. 2 g

(vii) Which of the following solution is more dilute?
A. 1M B. 2M C. 0.1M D. 0.009M

(viii) A solution of NaOH contains 20g of this compound in 2dm^3 of solution. What is the molarity of this solution?
A. 2M B. 1M C. 0 25M D. 0.5M

(ix) Commercially concentrated HCl is
A. 0.4 B. 0.5 C. 0 8 D. 0.37

(x) If one mole of Na contains x atoms of sodium, what is the number of moles contained in 46g of sodium?

(xii) Brass is alloy of ____________.
A. Fe + Cr B. Ag + Cu C. Cu + Zn D. Cu + Cr

Time allowed: 2:40 hours Total Marks: 53

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet–B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. Differentiate between saturated and unsaturated solution?

ii. Write the difference between a dilute and a concentrated solution?

iii. Can you call collide a solution?

iv. Calculate the new molarity when 100 cm^3 of water is added to 100 cm^3 of 0.5M HCl.

v. Benzene is a common organic solvent. Its use is now restricted because this can cause cancer. The recommended limit of exposure to benzene is 0.32 mg per dm^3 of air. Calculate the molarity of this solution.

vi. Define Saturated, unsaturated, supersaturated solutions and dilution of solution.

vii. How will you know whether a solution is saturated or supersaturated?

viii. If you read label on the bottle of concentrated H_2SO_4 you will notice 98% H_2SO_4 by mass and also 18M H_2SO_4. What does 18M stands for?

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. Define molarity and give its units.

ii. How does the change of temperature affect the solubility of certain compounds in water?

iii. What are true solutions explain with example?

iv. 100 cm^3 of NaOH solution was heated to complete dryness, 1.5 g residue left behind. What was the molarity of the solution.

v. Explain the solubility and solute-solvent interactions with the example of hydrogen bonding between water and methanol molecules.

vi. Compare the true solution, colloidal solution and suspension in all respects.

vii. Define concentration. Write the names of the concentration units of a solution?

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 a. Explain why CH_3OH is soluble in water but C_6H_6 is not.

b. What is the molarity of a solution prepared by dissolving 1.25 g of HCl gas into enough water to make 30 cm^3 of solution.

Q.5 a. Explain the effect of temperature on solubility with the help of solubility curves?

b. A solution of $Ca(OH)_2$ is prepared by dissolving 5.2 mg of $Ca(OH)_2$ to a total volume of 1000 cm^3. Calculate the molarity of this solution.

Q.6 a. Give the detail of percentage % compositions with specific examples.

b. Define solubility. How does nature of solute and solvent determine the extent of dissolution?

SOLUTION OF GUESS PAPER & MODEL PAPER # 6
(Reduced Syllabus)

SECTION- A (MCQs)

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. **($6 \times 3 = 18$)**

i. Differentiate between saturated and unsaturated solution?

Ans: Saturated solution: The solution which cannot dissolve more solute at a particular temperature is called a saturated solution.

Unsaturated solution: A solution which can dissolve more of the solute at a given temperature is called an unsaturated solution.

ii. Write the difference between a dilute and a concentrated solution?

Ans: Difference between a dilute and a concentrated solution:

A dilute solution is that whose concentration is relatively low while a concentrated solution is that solution whose concentration is relatively high.

iii. Can you call collide a solution?

Ans: Yes, collides are heterogeneous solution.

Examples: i. Starch solution ii. White of an egg.
iii. Gelatin, glue, gums Milk, rubber, fog, dust in the air, jellies, paints, blood and starch in water.

iv. Calculate the new molarity when 100 cm^3 of water is added to 100 cm^3 of 0.5M HCl.

Solution: M_1 = new molarity = ?
V_1 = volume of solution = 100 + 100 = 200 cm^3
M_2 = Given molarity of HCl solution = 0.5 M
V_2 = volume of HCl = 100 cm^3

$$M_1V_1 = M_2V_2$$

$$M_1 = \frac{M_2V_2}{V_1} \Rightarrow M_1 = \frac{0.5 \times 100}{200} = \frac{0.5}{2} = 0.25\ M$$

v. Benzene is a common organic solvent. Its use is now restricted because this can cause cancer. The recommended limit of exposure to benzene is 0.32 mg per dm^3 of air. Calculate the molarity of this solution.

Solution: Mass of solute C_6H_6 = 0.32 mg = $\frac{0.32}{1000}$ = 0.00032g
Molar mass of solute C_6H_6 = $6 \times 12 + 1 \times 6 = 72 + 6 = 78$ g
Volume of solution = 1 dm^3 = 1000 cm^3
Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{0.00032}{78} \times \frac{1000}{1000} = \frac{0.00032}{78} = 0.000004\ \text{M}$$

$$\text{Molarity(M)} = 4 \times 10^{-6}\ \text{M}$$

vi. Define Saturated, unsaturated, supersaturated solutions and dilution of solution.

Ans: Saturated solution: The solution which cannot dissolve more solute at a particular temperature is called a saturated solution.

Unsaturated solution: A solution which can dissolve more of the solute at a given temperature is called an unsaturated solution.

Supersaturated solution: A solution that contains more of the solute than is contained in the saturated solution is called supersaturated solution.

Examples: Take 100g water in a beaker. Add a tea spoon of sugar in it. Stirr it, the sugar will dissolve. Repeat the process the added sugar will again dissolve in it. This means solution is unsaturated.

Go on adding sugar in the above solution till it starts settling down at the bottom of the beaker at a particular temperature. This means solution is saturated.

vii. How will you know whether a solution is saturated or supersaturated?

Ans: A supersaturated solution is not stable in the presence of crystals of solute. If you add a crystal of sodium thiosulphate to its saturated solution, it will simply drop to the bottom, without dissolving, But if you add a crystal of sodium thiosulphate to a supersaturated solution of sodium thiosulphate (see figure a), crystallization will start. When crystallization has finished, you will have a saturated solution in presence of sodium thiosulphate crystals. Figure (b) shows the crystallization from a supersaturated solution of sodium thiosulphate.

(a)

(b)

Crystallization from a supersaturated solution of sodium thiosulphate

viii. If you read label on the bottle of concentrated H_2SO_4 you will notice 98% H_2SO_4 by mass and also 18M H_2SO_4. What does 18M stands for?

Ans: This means there are 18 moles of H_2SO_4 in each dm^3 of solution. Similarly conc. HCl is 37% and 12.1 M HCl. This means there are 12.1 moles of HCl in each dm^3 of solution.

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. **(5 × 3 = 15)**

i. Define molarity and give its units.

Ans: Molarity (M):

Molarity is the concentration unit in which amount of solute is expressed in moles and quantity of solution in dm^3.

"Molarity is defined as the number of moles of solute dissolved per dm^3 of solution".

Formula of Molarity:

$$\text{Molarity(M)} = \frac{\text{Number of moles of solute}}{\text{Volume of solution in dm}^3}$$

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Volume of solution in dm}^3}$$

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

Units of Molarity: mol dm^{-3}

ii. How does the change of temperature affect the solubility of certain compounds in water?

Ans: Effect of temperature on solubility:

Change of temperature can change the solubility of a solute in a solvent. It may increase or decrease. Consider the compounds such as KNO_3, KCl, $AgNO_3$ and KI etc. Heat is absorbed when solutions of these substances are formed in water.

When these substances are dissolved, the vessel cools down. The reason is that during dissolution the heat of solvent and the vessel is taken up in the process of solution formation. Whenever temperature of such solution is increased, solubilities of solutes increase.

A few substances like lithium carbonate $(Li_2CO)_3$, calcium hydroxide $Ca(OH)_2$, etc., dissolve in water with the evolution of heat. Their solubilities decrease by increase of temperature. Such solutions are warmed up along with the vessel. The solubility of NaCl in water is least affected by change of temperature. The reason is that a very small amount of heat is absorbed during the solution formation.

iii. What are true solutions explain with example?

Ans: True solution:

Properties of true solution:

Sizes of these particles are between **0.1-1 nm (1 nm = 10^{-9}m)**. Therefore, these particles cannot be seen by the naked eye, ordinary microscope and electron microscope. These particles can pass through ordinary and ultra filter papers. A solution is unable to scatter light

Examples: Dissolution of sodium Chloride or copper Sulphate in water

iv. **100 cm^3 of NaOH solution was heated to complete dryness, 1.5 g residue left behind. What was the molarity of the solution.**

Solution: Mass of solute NaOH = 1.5 g ; Molar mass of solute NaOH = 23 + 16 + 1 = 40 g
Volume of solution = 100 cm^3 ; Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{1.5}{40} \times \frac{1000}{100} = \frac{15}{40} = 0.375\ \text{M}$$

v. **Explain the solubility and solute-solvent interactions with the example of hydrogen bonding between water and methanol molecules.**

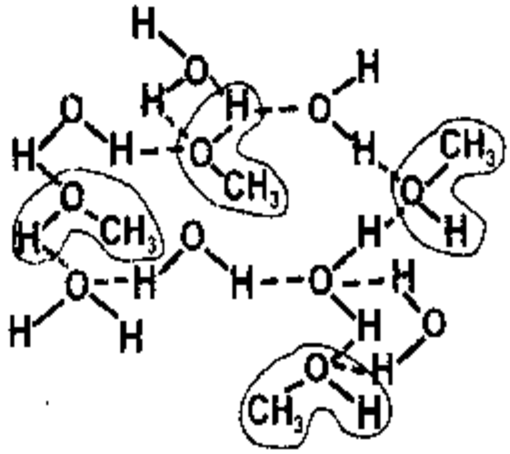
OR

Explain how methanol and water are miscible?

Ans: Solubility and solute-solvent interactions:

"like dissolves like" is a guiding rule when considering the solubility of substances.

Methanol readily dissolves in water. Water molecules are polar. Two H-atoms bonded to an O-atom are slightly positively charged and O-atom has a slightly negative charge. Water molecules form hydrogen bonds with one another. Methanol molecules are also polar and exhibit hydrogen-bonding. This means water and methanol molecules have similar structures and intermolecular forces. They can form hydrogen bonds with each other. Thus methanol and water are miscible.



H-bonding between water and methanol molecules

Similarly glucose, whose molecule has many –O-H bonds, is very soluble in water.

vi. **Compare the true solution, colloidal solution and suspension in all respects.**

Ans: Comparison of properties of solutions, suspension and colloids:

No.	True Solutions	Suspensions	Colloids
1.	Homogeneous	Heterogeneous	Heterogeneous
2.	Particles size vary from 0.1 to 1 nm	Particles size is greater than 10^3 nm	Particles size vary from 1 to 10^3 nm
3.	Particles are invisible by naked eye, ordinary microscope as well as electron microscope	Particles are visible by naked eye	Particles are invisible by naked eye and in ordinary microscope but visible under electron microscope
4.	Particles can pass through ordinary as well as ultra filter paper	Particles cannot pass through ordinary as well as ultra filter paper	Particles can pass through ordinary filter paper but cannot pass through ultra filter paper
5.	Cannot scatter light	Scatter light	Scatter light
6.	**Examples:** i. Dissolution of sodium Chloride or copper Sulphate in water.	**Examples:** i. A mixture of black board chalk in water. ii. smoke mud, emulsion, clay, sand in water	**Examples:** i. Starch solution ii. White of an egg. iii. Gelatin, glue, gums Milk, rubber, fog, dust in the air

vii. Define concentration. Write the names of the concentration units of a solution?

Ans: Concentration:

The quantity of a solute present in a given amount of solvent or solution is called concentration of solution.

Concentration units of a solution: **i.** Percentage % **ii.** Molarity M

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. **($2 \times 10 = 20$)**

Q.4 a. Explain why CH_3OH is soluble in water but C_6H_6 is not.

Ans: Methanol (CH_3OH) is soluble in water due to hydrogen bonding.

According to the general principle "Like dissolve like". As water is polar solvent and Benzene is non polar, therefore, benzene will not dissolve in water.

b. What is the molarity of a solution prepared by dissolving 1.25 g of HCl gas into enough water to make 30 cm^3 of solution.

Solution: Mass of solute HCl = 1.25 g ; Molar mass of solute HCl = 1 + 35.5 = 36.5 g

Volume of solution = 30 cm^3 ; Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{1.25}{36.5} \times \frac{1000}{30} = \frac{1250}{1095} = 1.14\ \text{M}$$

Q.5 a. Explain the effect of temperature on solubility with the help of solubility curves?

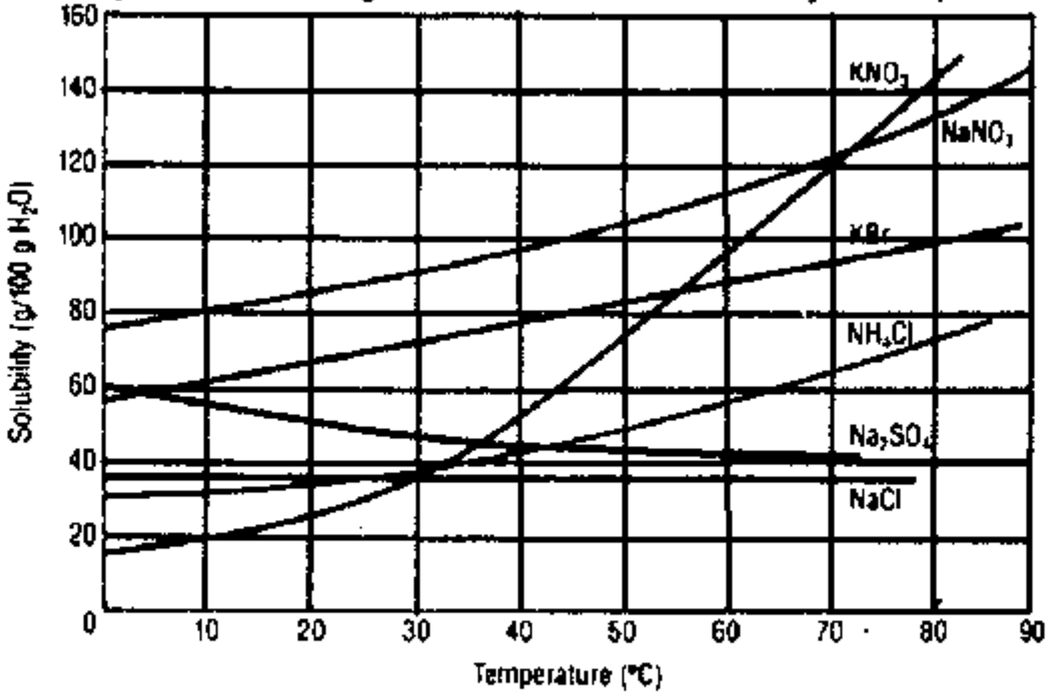
OR

Explain the variation of solubility with temperature?

Ans: Variation of solubility with temperature:

If we add 34.7 g of KCl to 100 g of water at 20 °C, it will dissolve. If we add more than 34.7 g of KCl at 20 °C, it will not dissolve. However, if we increase temperature it will readily dissolve. Keep on adding more KCl and increase temperature.

We will observe that 56.7 g of KCl dissolves in 100 g of water at 100°C, when temperature of this solution is decreased to 20 °C, 22.0 g of KCl will crystallize out. This means temperature has a profound effect on the solubility of a substance. Figure shows the variation of solubility with temperature.



Variation of solubility with temperature

Solubility of ionic compounds:

For instance, solubility of Na_2SO_4 decreases with increase in temperature.

Solubility of gases with increasing temperature:

Heat water in a beaker, you will see small bubbles form at the side of the beaker before the water boils. These bubbles are composed of air. Since, air is less soluble in hot water than in cold water, air comes out of water in the form of bubbles. This means solubility of air in water decreases with increasing temperature. You might have observed in a home aquarium, that the fish shows signs of stress on a hot day. This is because less oxygen from air dissolves in the warm water

b. A solution of $Ca(OH)_2$ is prepared by dissolving 5.2 mg of $Ca(OH)_2$ to a total volume of 1000 cm^3. Calculate the molarity of this solution.

Solution: Mass of solute $Ca(OH)_2$ = 5.2 mg = $\frac{5.2}{1000}$ = 0.0052 g

Molar mass of solute $Ca(OH)_2$ = 40 + 2 × 16 + 2 × 1 = 40 + 32 +2 = 74 g

Volume of solution = 1000 cm^3 ; Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{0.0052}{74} \times \frac{1000}{1000} = \frac{5.2}{74000}$$

$$\text{Molarity(M)} = 0.00007 = 7 \times 10^{-5}\ \text{M}$$

Q.6 a. Give the detail of percentage % compositions with specific examples.

Ans: Percentage % composition:

By the percentage of a solution we mean the mass or volume of solute dissolved in 100 g or 100 cm^3 of solution.

Examples: Commercially available solutions that contain the maximum available concentration of solutes are referred as concentrated solutions. For example commercially available concentrated H_2SO_4 is 98% and conc HCl is 37%. This means concentrated H_2SO_4 contains 98g H_2SO_4 in 100g of solution and concentrated solution of HCl contains 37g HCl in 100g of solution.

Thus percentage is the unit of concentration that specifies the quantity of solute in 100 parts of solution. Quantity of solute and solvent can be expressed by mass in grams or volume in cm^3.

10 % m/m concentration:

For example, if you dissolve 10 g NaCl in 90 g water to make 100 g of solution, the concentration of solution will be **10% m/m.**

10 % v/v concentration:

But if you dissolve 10 cm^3 of alcohol in sufficient water to make **100 cm^3** of solution, the solution will be **10% v/v.**

10% m/v concentration:

If you dissolve 10g NaCl in sufficient water to make **100 cm^3** solution, the resulting solution will be **10% m/v.**

10% v/m concentration:

If you dissolve **10 cm^3** of alcohol in water to make 100 g of solution, the solution will be **10% v/m.** Therefore, there are four ways to express percentage of solutions.

b. Define solubility. How does nature of solute and solvent determine the extent of dissolution?

Ans: Solubility:

The amount of solute that dissolves in 100g of a solvent at a particular temperature is called its solubility

OR

Solubility of a substance in a particular solvent at a definite temperature is the maximum amount of the solute in grams that can dissolve in 100 grams of the solvent to form a saturated solution.

Examples: **The solubility of sodium chloride (NaCl) in 100 g of water at 100°C is 39.12 g.**

The solubility of silver chloride (AgCl) in 100 g water at 100°C is 0.0021 g.

Obviously, NaCl is much more soluble in water than AgCl.

Nature of solute and solvent determine the extent of dissolution:

Solubility of ionic and polar compounds:

It means that the solubility of a solute in a solvent depends upon the nature of both. The ionic and polar compounds like NaCl and HCl are more soluble in water than non-polar covalent compounds like CS_2 and CCl_4.

Solubility of non- polar covalent compounds:

The non-polar covalent compounds freely dissolve in non-polar covalent solvents like CCl_4 and benzene, etc. Iodine is freely soluble in CCl_4 because both the solute and the solvent are non-polar and their intermolecular attractions are of the same order. Waxes and fats are soluble in benzene and not in water.

The molecules of non-polar covalent substances like benzene (C_6H_6), carbon tetrachloride (CCl_4) have their **dipole moments very close to zero.** So, they have weak van der Waal's forces.

Non-polar substances like C_6H_6 and CCl_4 are immiscible in polar solvents like H_2O:

On the other hand, polar covalent substances like water (H_2O) have permanent dipoles. These permanent dipoles are responsible for strong intermolecular attractions among water molecules. So, non-polar substances like C_6H_6 and CCl_4 are immiscible in polar solvents like H_2O This is due to the reason that the attraction of a water molecule is much greater for one another than the attraction between water and benzene (C_6H_6) molecules.

IMPORTANT SHORT QUESTION ANSWERS (Reduced Syllabus)

SELF ASSESSMENT EXERCISE 6.1

The maximum amount of sodium acetate that dissolves in 100g of water at 0°C is 119g and 170g at 100°C.

(a) If you add 170g of sodium acetate in 100g of water at 0°C, how much will dissolve?
(b) Is the solution saturated unsaturated or supersaturated?
(c) If the solution in heated to 100°C, is the solution now saturated, unsaturated or supersaturated.
(d) If the solution is cooled back to 0°C and no crystals appear. Is the solution saturated, unsaturated or supersaturated?

Solution: **(a)** 119 g. **(b)** Saturated **(c)** Supersaturated. **(d)** Unsaturated

SELF ASSESSMENT EXERCISE 6.3

1. **Write four ways to express percentage of solutions.**
2. **A saline solution is administered intravenously to a person suffering from severe dehydration. This is labeled as 0.85% m/v of NaCl. What does it mean?**

Solution: **1. Write four ways to express percentage of solutions.**

Ans: Four ways to express percentage of solutions:

(i) Mass by mass percent **(ii)** Mass by volume percent
(iii) Volume by mass percent **(iv)** Volume by volume percent

(i) Mass by mass percent: It is the mass of the solute dissolved per 100 parts by mass of solution.

$$\%\text{ by mass}\left(\frac{M}{M}\right) = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

(ii) Mass by volume percent: It is the mass of the solute dissolved per 100 parts by volume of solution.

$$\%\text{ of solution}\left(\frac{M}{V}\right) = \frac{\text{Mass of solute}}{\text{Volume of solution}} \times 100$$

(iv) Volume by volume percent: It is the volume of the solute dissolved per 100 parts by volume of solution.

$$\% \text{ of solution by } \left(\frac{V}{V}\right) = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

2. A saline solution is administered intravenously to a person suffering from severe dehydration. This is labeled as 0.85% m/v of NaCl. What does it mean?

Ans: 0.85% m/v concentration: If you dissolve 0.85 g NaCl in sufficient water to make 100 cm^3 solution, the resulting solution will be 0.85% m/v.

SELF ASSESSMENT EXERCISE 6.4

Potassium chlorate ($KClO_3$) is a white solid. It is used in making matches and dyes. Calculate the molarity of solution that contains.

(a) 1.5 moles of this compound dissolved in 250 cm^3 of solution

(b) 75 g of this compound dissolved to produce 1.25 dm^3 of solution.

(c) What is the molarity of a 50 cm^3 sample of potassium chlorate solution that yields 0.25 g residue after evaporation of the water.

Solution:

(a) 1.5 moles of this compound dissolved in 250 cm^3 of solution

Ans: Volume of solution in $dm^3 = \frac{250}{1000} dm^3 = 0.25 \ dm^3$

Number of moles = 1.5 moles ; Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Number of moles}}{\text{Volume of solution in dm}^3}$$

$$\text{Molarity(M)} = \frac{1.5}{0.25} \quad \Rightarrow \quad \text{Molarity(M)} = 6 \text{ M}$$

(b) 75 g of this compound dissolved to produce 1.25 dm^3 of solution.

Ans: Mass of solute $KClO_3$ = 75 g ; Molar mass of solute $KClO_3$ = 39 + 35.5 + 48 = 122.5 g

Volume of solution = 1.25dm^3 = 1.25 × 1000 = 1250 cm^3

Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{75}{122.5} \times \frac{1000}{1250} = \frac{75000}{153125} \quad \Rightarrow \quad \text{Molarity(M)} = 0.49 \text{ M}$$

(c) What is the molarity of a 50 cm^3 sample of potassium chlorate solution that yields 0.25 g residue after evaporation of the water.

Ans: Mass of solute $KClO_3$ = 0.25 g ; Molar mass of solute $KClO_3$ = 39 + 35.5 + 48 = 122.5 g

Volume of solution = 50 cm^3 ; Molarity (M) = ?

$$\text{Molarity(M)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{0.25}{122.5} \times \frac{1000}{50} = \frac{250}{6125} \quad \Rightarrow \quad \text{Molarity(M)} = 0.04 \text{ M}$$

SELF ASSESSMENT EXERCISE 6.5

1. Sodium hydroxide solutions are used to neutralize acids and in the preparation of soaps and rayon. If you dissolve 25g of NaOH to make 1 dm^3 of solution, What is the molarity of

3. A solution is prepared by dissolving 10 g of hemoglobin in enough water to make up 1 dm^3 in volume. Calculate molarity of this solution. Molar mass of hemoglobin is 6.51×10^4 g/mole.

Solution:

1. **Sodium hydroxide solutions are used to neutralize acids and in the preparation of soaps and rayon. If you dissolve 25g of NaOH to make 1 dm^3 of solution, What is the molarity of this solution?**

Solution: Mass of solute NaOH = 25 g ; Molar mass of solute NaOH = 23 + 16 + 1 = 40 g
Volume of solution = 1dm^3 = 1000 cm^3 ; Molarity (M) = ?

$$\text{Molarity}(M) = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity}(M) = \frac{25}{40} \times \frac{1000}{1000} = \frac{25}{40} \quad \Rightarrow \quad \text{Molarity}(M) = 0.625\ M$$

2. **A solution of NaOH has concentration 1.2M. Calculate the mass of NaOH in g/dm^3 in this solution.**

Solution:

Volume of solution = 1 dm^3 = 1000 cm^3
Molarity (M) = 1.2 M
Molar mass of solute NaOH = 23 + 16 + 1 = 40 g
Mass of solute = ?

$$\text{Molarity}(M) = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$1.2 = \frac{\text{Mass of solute}}{40} \times \frac{1000}{1000}$$

$$1.2 = \frac{\text{Mass of solute}}{40}$$

$$\text{Mass of solute} = 1.2 \times 40 = 48\ g/dm^3$$

3. **A solution is prepared by dissolving 10 g of hemoglobin in enough water to make up 1 dm^3 in volume. Calculate molarity of this solution. Molar mass of hemoglobin is 6.51×10^4 g/mole.**

Solution:

Mass of solute hemoglobin = 10 g
Molar mass of solute hemoglobin = 6.51×10^4 g/mole.
Volume of solution = 1dm^3 = 1000 cm^3
Molarity (M) = ?

$$\text{Molarity}(M) = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity}(M) = \frac{10}{6.51 \times 10^4} \times \frac{1000}{1000}$$

$$\text{Molarity}(M) = \frac{10}{6.51 \times 10^4}$$

$$\text{Molarity}(M) = 1.54 \ldots M$$

SELF ASSESSMENT EXERCISE 6.8

1. **Sodium Chloride and glucose both are soluble in water. But the Solubility of NaCl is greater than glucose. Explain why?**
2. **In which liquid of each of the following pairs you would expect KCl, an ionic solid, to be more soluble. (a) H_2O or CCl_4 (b) CH_3OH or Benzene**
3. **Which of the following pairs of liquids are miscible?**
   (a) Water and benzene (b) Benzene and CCl_4 (c) An oil and benzene

Solution:

1. **Sodium Chloride and glucose both are soluble in water. But the Solubility of NaCl is greater than glucose. Explain why?**

Ans: The solubility of sodium chloride in water is far greater than that of glucose. This is due to the fact that the attraction of sodium (Na^+) and chloride (Cl^-) ions with water is greater than that of glucose molecules with water.

2. **In which liquid of each of the following pairs you would expect KCl, an ionic solid, to be more soluble. (a) H_2O or CCl_4 (b) CH_3OH or Benzene**

Ans: a. H_2O due to its high dielectric constant (polar nature), breaks the crystal lattice in between K^{+1} and Cl^{-1}. Therefore KCl dissolves in H_2O.

According to the general principle "Like dissolve like" As water is polar solvent and carbon tetra chloride (CCl_4) is non polar, therefore, KCl will not dissolve in CCl_4

b. KCl is soluble in CH_3OH (Methanol) due to hydrogen bonding.

According to the general principle "Like dissolve like". As KCl is ionic and benzene (C_6H_6) is non polar. therefore, benzene will not dissolve in KCl.

3. **Which of the following pairs of liquids are miscible?**
   (a) Water and benzene (b) Benzene and CCl_4 (c) An oil and benzene

Ans: (a) Water and benzene:

According to the general principle "Like dissolve like". As water is polar solvent and benzene (C_6H_6) is non polar, therefore, benzene will not dissolve in water. Therefore water and benzene are not miscible.

(b) Benzene and CCl_4:

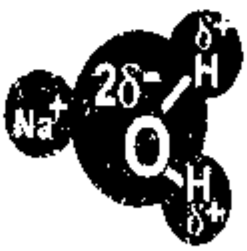
According to the general principle "Like dissolve like . As benzene is non polar solvent and CCl_4 is also non polar, therefore, benzene and CCl_4 are miscible

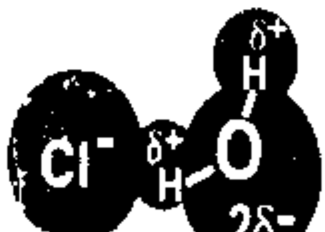
(c) An oil and benzene:

According to the general principle "Like dissolve like". As benzene is non polar solvent and oil is also non polar therefore, benzene and oil are miscible.

Q3. Explain the solubility and solute-solvent interactions with the example of attraction of Na^+ and Cl^- ions for water molecules. OR
Explain the attraction of Na^+ and Cl^- ions for water molecules?

Ans: When we place a crystal of sodium chloride in water it dissolves. Sodium chloride in an ionic compound. The negative end of water molecules is attracted to sodium ions and the positive end of water molecules is attracted to chloride ions. These attractive forces are strong enough to overcome the strong attractions that exist between ions in NaCl. Thus sodium chloride dissolves readily.





1: **A 10.0 g of solid solute is placed in 100 g of water at 20 °C and all of it dissolves. Then another 4.0 g of the solute is added at 20 °C and all of it dissolves.**

(a) **Is the first solution saturated, unsaturated or supersaturated?**

(b) **Is it possible to tell from this information that the final solution is unsaturated or saturated?**

Ans: **(a)** Unsaturated.

(b) Unsaturated. Because a solution which can dissolve more of the solute at a given temperature is called an unsaturated solution.

2: **What should you do to change:**

(a) **a saturated solution to an unsaturated solution.**

(b) **an unsaturated solution to a saturated solution.**

Ans: **(a)** **i.** Add more solvent. **ii.** increase temperature.

(b) Keep adding solute until the solvent cannot dissolve any more at that given temperature.

9: **A patient in a hospital is often administered an intravenous (IV) drip containing an aqueous solution. This aqueous solution contains 0.85% (mass by volume) of sodium chloride or 5% (mass by volume) of glucose. Calculate the molarity of both these solutions.**

Solution:

Case i. **0.85% m/v of NaCl**

Mass of solute = 0.85 g

Molar mass of NaCl = 23 + 35.5 = 58.5 g / mol

$$\text{Molarity(M)} = \frac{\text{Mass}}{\text{Molar mass}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{0.85}{58.5} \times \frac{1000}{100}$$

$$\text{Molarity(M)} = 0.145\ \text{M}$$

Case ii. **5% m/v of glucose ($C_6H_{12}O_6$)**

Molarity solute = 5 g

Volume of solution = 100 cm^3

Molar mass of glucose = 12 × 6 + 1 × 12 + 16 × 6 = 180 g/mol

$$\text{Molarity(M)} = \frac{\text{Mass}}{\text{Molar mass}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

$$\text{Molarity(M)} = \frac{5}{180} \times \frac{1000}{100}$$

$$\text{Molarity(M)} = 0.278\text{M}$$

GUESS PAPER & MODEL PAPER # 7
BASED ON CHAPTER # 7 (Reduced Syllabus)
ELECTROCHEMISTRY

CHAPTER 7: ELECTROCHEMISTRY

Oxidation, Reduction, oxidation reduction in term of loss or gain of electrons, oxidation state and rules for assigning oxidation state and oxidation number, oxidizing and reducing agent, Electrochemical cell, nature of electrochemical process, concept of electrolyte, electrolytic cell, use of electrolytic cell, galvanic cell (Daniel cell), Reaction in Daniel cell, Zinc electroplating, chromium electroplating, corrosion and its prevention, rusting of iron, corrosion of Aluminum, cathode protection.

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) In which of the following changes the nitrogen atom is reduced.
A. N_2 to NO B. N_2 to NO_2 C. N_2 to NH_3 D. N_2 to HNO_3

(ii) Which of the following changes reaction is an example of oxidation?
A. Chlorine molecule to chloride ion B Silver atoms to silver (I) ion
C. Oxygen molecule to oxide ion D. Iron (III) ion to iron(II) ion

(iii) On an industrial scale, sodium metal is prepared by electrolysis of fused
A. $NaOH$ B. $NaCl$ C. NaO D. NH_3

(iv) Corrosion can be prevented by
A. Alloying B. Tinning C. Galvanizing D. all of above

(v) The oxidation state of Cr in $K_2Cr_2O_7$ is
A. 12 B. 6 C. 3 D. Zero

(vi) Which of the following statement is not correct about the galvanic cell?
A. Cations are reduced at cathode B. Anions are oxidized at anode
C. Electrons flow from cathode to anode D. Oxidation occurs at anode

(vii) An electrolytic cell uses electrical energy to drive
A. chemical reaction B. physical reaction
C. no reaction D. biochemical reaction

(viii) An electrochemical cell is also called
A. battery cell B. galvanic cell C. cell D. chargeable cell

(ix) Galvanizing is __________.
A. Coating with Sn B. Coating with Zn
C. Coating with Cr D. Coating with Cu

(xi) Elements in free state have ______________ oxidation number:
A. 1 B. 0 C. 2 D. 3

(xii) Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:
A. Cl_2 B. H_2 C. O_3 D. O_2

Time allowed: 2:40 hours Total Marks: 53

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet-B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. What is oxidation state?
ii. What is the oxidation number of Cr in chromic acid (H_2CrO_4)?
iii. List the possible uses of electrolytic cells?
iv. Do you think H in HCl is oxidized and Cl is reduced?
v. Compare the process of oxidation and reduction?
vi. Explain one example from daily life which involves oxidation-reduction reaction?
vii. Find the oxidation state of S in the following compounds.
(a) H_2S (b) H_2SO_3 (c) $Na_2S_2O_3$
viii. Describe the rules for assigning oxidation states or numbers?

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. (5 × 3 = 15)

i. Consider the reaction. $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
a. Which substance in this reaction loses oxygen?
b. Which substance gains oxygen?
c. Which substance is oxidized?
d. which substance is reduced in the above reaction?
ii. What is oxy-acetylene flame?
iii. Consider the following reaction in which acetylene (C_2H_2) burns.
$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$
a. Which substance loses hydrogen?
b. Which substance gains hydrogen?
iv. Define oxidation and reduction in terms of loss or gain of electrons?
v. Define oxidizing and reducing agent. Point out the oxidizing and reducing agent in the reactions with the help of examples?
vi. Define electrochemical cells. Describe their different types?
vii. Describe the nature of electrochemical process?

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 State the substances which are oxidized or reduced. Give reason for your answer.

(a) $N_2 + 3H_2 \longrightarrow NH_3$

(b) $CO_2 + 2Mg \longrightarrow 2MgO + C$

(c) $Mg + H_2O \longrightarrow MgO + H_2$

Q.5 a. Sketch a Daniel Cell, labeling the cathode, anode, and the direction of flow of the electrons.
b. Compare the effects of Al_2O_3 and Fe_2O_3 formation on their parent metals and cite examples from daily life.

Q.6 Electrolysis has a major role in electrochemical industries.
(a) Sketch an electrolytic cell, label the anode and cathode and indicate the direction of electron transfer.
(b) Describe the nature of electrochemical process.
(c) Distinguish between electrolytic and voltaic cell.

SOLUTION OF GUESS PAPER & MODEL PAPER # 7
(Reduced Syllabus)

SECTION- A (MCQs)

i. C	ii. B	iii. B	iv. D	v. B	vi. C
vii. A	viii. B	ix. B	x. D	xi. B	xii. B

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. **(6 × 3 = 18)**

i. What is oxidation state?

Ans: Oxidation states or oxidation number:

Oxidation state or oxidation number is defined as the number of charges an atom will have in a molecule or a compound.

The elements that show an increase in oxidation number are oxidized. The elements that show a decrease in oxidation number are reduced

ii. What is the oxidation number of Cr in chromic acid (H_2CrO_4)?

Solution: Oxidation number of Cr in chromic acid (H_2CrO_4):

The sum of oxidation numbers must be zero.

2[O.N of H] + [O.N of Cr] + 4 [O.N of O] = 0 ⇒ 2[+1] + [O.N of Cr] + 4 [-2] = 0

+2 + [O.N of Cr] - 8 = 0 ⇒ [O.N of Cr] - 6 = 0 ⇒ [O.N of Cr] = +6

Thus oxidation state for Cr in H_2CrO_4 is +6.

iii. List the possible uses of electrolytic cells?

Ans: Uses of electrolytic cells:

Possible uses of electrolytic cells are as follows.

(a) Down's Cell is used for the commercial preparation of sodium metal. It produces chlorine gas as by product.
(b) Nelson's Cell is used for the commercial preparation of sodium hydroxide. It also produces chlorine and hydrogen gas as by product.
(c) Electrolytic cells are used for the commercial preparation of calcium and magnesium metals.
(d) It is used to produce aluminum metal commercially
(e) It is used for the purification of copper.
(f) Electrolytic cells are used to electroplate metals such as tin, silver, nickel etc on steel.
(g) Electrolytic cells are used to prepare anodized aluminum. Anodized aluminum can absorb dyes. Dyeing of anodized aluminum can produce metallic red, metallic blue or other metallic colours on the metal surface.

iv. Do you think H in HCl is oxidized and Cl is reduced?

Ans:

Oxidation state of Cl in HCl is -1 (gain of one electron) Hence Cl is reduced.

v. Compare the process of oxidation and reduction?

Ans: Comparasion of oxidation and reduction processes:

Oxidation	Reduction
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electrons
increase in oxidation number	Decrease in oxidation number

vi. Explain one example from daily life which involves oxidation-reduction reaction?

Ans: Redox in photography:

A photographic film is basically an emulsion of silver bromide, (AgBr) in gelatin. When the film is exposed to light, Silver bromide granules become activated. This activation depends on the intensity of the light falling upon them. When exposed film is placed in the developer solution that is actually a reducing agent. Hydroquinone which is a mild reducing agent is used as developer. In hydroquinone the activated granules of silver bromide are reduced to black metallic silver. Reduced silver atoms form image.

$$Ag^{+} + 1e^{-} \longrightarrow Ag(s)$$

Inactivated silver bromide is removed from the film by using a solvent called a fixer. Sodium thiosulphate is used for this purpose. The areas of the film exposed to the light appear darkest because they have the highest concentration of metallic Silver. Thus photography involves oxidation-reduction reaction.

vii. Find the oxidation state of S in the following compounds.

(a) H_2S (b) H_2SO_3 (c) $Na_2S_2O_3$

Ans: (a) Oxidation number of S in H_2S:

The sum of oxidation numbers must be zero.

[O.N of S] + 2[O.N of H] = 0 ⇒ [O.N of S] + 2[+1] = 0 ⇒ [O.N of S] + 2 = 0
[O.N of S] = -2 ; Thus oxidation state for S in H_2S is -2.

(b) Oxidation number of S in H_2SO_3:

The sum of oxidation numbers must be zero.

2[O.N of H] + [O.N of S] + 3 [O.N of O] = 0 ⇒ 2[+1] + [O.N of S] + 3 [-2] = 0
+2 + [O.N of S] -6 = 0 ⇒ [O.N of S] - 4 = 0 ⇒ [O.N of S] = +4
Thus oxidation state for S in H_2SO_3 is +4.

(c) Oxidation number of S in $Na_2S_2O_3$:

The sum of oxidation numbers must be zero.

2[O.N of Na] + 2[O.N of S] + 3[O.N of O] = 0 ⇒ 2[+1] + 2[O.N of S] + 3[-2] = 0
+2 + 2[O.N of S] -6 = 0 ⇒ 2[O.N of S] - 4 = 0 ⇒ 2[O.N of S] = +4

$$\frac{2[\text{O.N of S}]}{2} = \frac{+4}{2} \Rightarrow [\text{O.N of S}] = +2$$

; Thus oxidation state for S in $Na_2S_2O_3$ is +2.

viii. Describe the rules for assigning oxidation states or numbers?

Ans: Rules for assigning oxidation states or numbers:

1. The oxidation state of any uncombined or free elements is always zero e.g., oxidation state of Zn, Na, H in H_2, S in S_8 etc is zero.
2. In simple ions, oxidation state is same as their charge e.g., oxidation state of Na in Na^{+1} and Ca in Ca^{+2} are +1 and +2 respectively.
3. In a complex ion the total sum of oxidation states of atoms is equal to the charge on their ion. e.g., in CO_3^{-2}, the sum of oxidation states of C and 3O atoms is -2. Similarly, in NH_4^{+1}, the sum of oxidation states of N and 4H atoms is +1.

Examples:

Elements	Oxidation State
Group-IA	+1
Group-IIA	+2
Group-IIIA	+3
H	+1 (except in metal hydrides where it is -1)
Group-VIIA	-1
O	-2 (except peroxides and in OF_2)

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. **(5 × 3 = 15)**

i. Consider the reaction. $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

- **a. Which substance in this reaction loses oxygen?**
- **b. Which substance gains oxygen?**
- **c. Which substance is oxidized.**
- **d. which substance is reduced in the above reaction?**

Solution: a. Fe_2O_3. b. CO. c. C d. Fe

ii. What is oxy-acetylene flame?

Ans: Oxy-acetylene flame:

Acetylene (C_2H_2) is commercially used for cutting and welding metals. When acetylene burns, it produces a very hot flame known as oxy-acetylene flame.

iii. Consider the following reaction in which acetylene (C_2H_2) burns.

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

- **a. Which substance loses hydrogen?** **b. Which substance gains hydrogen?**

Solution: a. C_2H_2 b. O

iv. Define oxidation and reduction in terms of loss or gain of electrons?

Ans: Oxidation: A process that involves the loss of electrons by an element is called oxidation.

Examples: For example, group IA and group IIA elements lose one and two electrons respectively to form cations. In doing so, these metals undergo oxidation

$$Na \longrightarrow Na^+ + e^-$$

$$Ca \longrightarrow Ca^{+2} + 2e^-$$

Other examples of oxidation are,

$$Fe \longrightarrow Fe^{+2} + 2e^-$$

$$Fe^{+2} \longrightarrow Fe^{+3} + e^-$$

$$Cu \longrightarrow Cu^{+2} + 2e^-$$

Reduction: A process that involves the gain of electrons by a substance is called reduction.

Examples: Elements of group VIA and VIIA gain one and two electrons respectively to form anions. They undergo reduction.

$$Cl + e^- \longrightarrow Cl$$

$$O + 2e^- \longrightarrow O^{-2}$$

$$S + 2e^- \longrightarrow S^{-2}$$

v. Define oxidizing and reducing agent. Point out the oxidizing and reducing agent in the

Reducing Agent:

A reducing agent is the reactant containing the element that is oxidized (loses electrons) in the chemical reaction.

Example: For example, in the reaction between sodium and chlorine to form sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Na is reducing agent hence being oxidised whereas Cl_2 is oxidizing agent hence being reduced

vi. Define electrochemical cells. Describe their different types?

Ans: Electrochemical cells:

Devices that convert chemical energy into electrical energy or vice versa are called electrochemical cells.

Types of electrochemical cell: Therefore, there are two types of electrochemical cells:

(i) Electrolytic Cells:

An electrochemical cell that uses electrical energy to drive a chemical reaction is called an electrolytic cell.

(ii) Galvanic Cells:

An electrochemical cell that converts chemical energy into electricity is called a galvanic cell.

vii. Describe the nature of electrochemical process?

Ans: Nature of electrochemical process:

Electrochemical processes are oxidation-reduction reactions in which chemical energy released by a spontaneous reaction is converted to electricity or in which electrical energy is used to drive a non-spontaneous reaction. Whether an electrochemical process releases or requires energy, it always involves the transfer of electrons from one substance to another. This means that this process always involves an oxidation-reduction or a redox reaction.

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. (2 × 10 = 20)

Q.4 State the substances which are oxidized or reduced. Give reason for your answer.

(a) $N_2 + 3H_2 \longrightarrow NH_3$

(b) $CO_2 + 2Mg \longrightarrow 2MgO + C$

(c) $Mg + H_2O \longrightarrow MgO + H_2$

(d) $H_2S + Cl_2 \longrightarrow 2HCl + S$

(e) $2NH_3 + 3CuO \longrightarrow 3Cu + N_2 + 3H_2O$

Solution:

a. $\overset{0}{N_2} + 3\overset{0}{H_2} \longrightarrow 2\overset{-3}{N}\overset{+3}{H_3}$

N-atoms undergo reduction, because there is a decrease in oxidation state of N (0 to -3).
H-atoms undergo oxidation, because there is a increase in oxidation state of H (0 to +3).

b. $\overset{0}{CO_2} + 2\overset{0}{Mg} \longrightarrow 2\overset{+2}{Mg}\overset{-2}{O} + \overset{0}{C}$

O-atoms undergo reduction, because there is a decrease in oxidation state of O (0 to -2).
Mg-atoms undergo oxidation, because there is a increase in oxidation state of Mg (0 to +2).

c. $\overset{0}{Mg} + \overset{2(+1)}{H_2}\overset{-2}{O} \longrightarrow \overset{+2}{Mg}\overset{-2}{O} + \overset{0}{H_2}$

H-atoms undergo reduction, because there is a decrease in oxidation state of H (+2 to 0).
Mg-atoms undergo oxidation, because there is a increase in oxidation state of Mg (0 to +2).

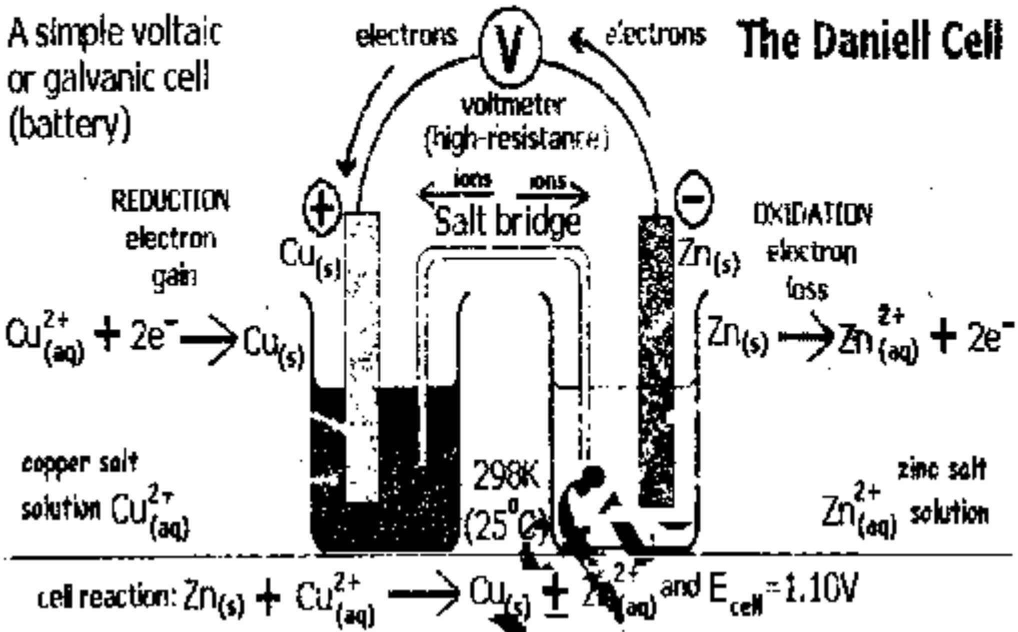
Cl-atoms undergo reduction, because there is a decrease in oxidation state of Cl (0 to -2).
S-atoms undergo oxidation, because there is a increase in oxidation state of S (-2 to 0).

e. $\overset{-3\ +3(1)}{2NH_3} + \overset{+2\ -2}{3CuO} \longrightarrow \overset{0}{3Cu} + \overset{0}{N_2} + \overset{2(+1)\ -2}{3H_2O}$

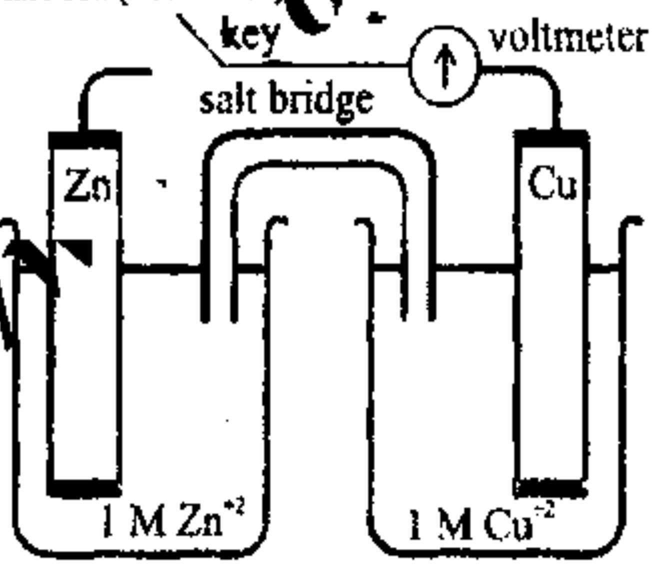
Cu-atoms undergo reduction, because there is a decrease in oxidation state of Cu (+2 to 0).
N-atoms undergo oxidation, because there is a increase in oxidation state of N (-3 to 0).

Q.5 a. Sketch a Daniel Cell, labeling the cathode, anode, and the direction of flow of the electrons.

Ans: Sketch of Daniel Cell:



Consider the following voltaic cell (Daniel cell):



(i) In which direction do electrons flow when the key is pressed?

Ans: Electrons leave the Zn-electrode and then move towards Cu- electrodes .

(ii) Oxidation takes place at which electrode?

Ans: Anode is Zn electrode. Oxidation takes place at zinc electrode.

(iii) Reduction takes place at which electrode?

Ans: Cathode is Cu electrode. reduction takes place at copper electrode.

(iv) In which half-cell do electrons enter the cell?

Ans: electrons enter the cell at copper electrode

(v) Electrons are consumed at which electrode?

Ans: Electrons are consumed at Cathode, right half cell

(vii) Which electrode will decrease in mass during operation of the cell?

Ans: Zinc electrode will decrease in mass because zinc atom loses electrons.

(viii) Which electrode will increase in mass during operation of the cell?

Ans: Copper (cathode) will increase in the mass.

(ix) Suggest a solution for the electrolyte.

Ans: 1M $ZnSO_4$ solution and 1M $CuSO_4$ solution

(x) In which direction do cations within the salt bridge move to maintain charge neutrality?

Ans: Negative ions move from cathodic compartment to anodic compartment through the salt bridge. In this way, the solutions of two half- cells remain neutral.

Salt bridge contains KCl. The cation (K^+) of salt bridge move towards copper electrode and Cl^- ions of salt bridge move towards zinc electrode.

(xi) Give balanced half-reactions and an overall cell reaction.

Ans: Right half-cell:

It contains a copper (Cu) metal strip dipped into 1 M solution of copper sulphate ($CuSO_4$) giving the following equilibrium:

$$Cu^{+2}_{(aq)} + 2e^- \rightleftharpoons Cu^{\circ}_{(s)}$$

Left half-cell:

It contains a zinc (Zn) metal strip dipped into 1 M solution of zinc sulphate ($ZnSO_4$) giving the following equilibrium:

$$Zn_{(s)} \rightleftharpoons Zn^{+2}_{(aq)} + 2e^{-2}$$

Redox reaction or overall voltaic cell reaction:

The sum of two half-cell reactions is called redox reaction or overall voltaic cell reaction.

$$Zn_{(S)} + Cu^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Cu_{(s)} \quad \text{---- (redox reaction)}$$

b. Compare the effects of Al_2O_3 and Fe_2O_3 formation on their parent metals and cite examples from daily life.

Ans: Corrosion of aluminum:

The corrosion resistance of aluminum is dependent upon a protective oxide film. This film is stable in aqueous media when the pH is between about 4.0 and 8.5. The oxide film is naturally self-renewing and accidental abrasion or other mechanical damage of the surface film is rapidly repaired.

However, this does not occur. This is because a tough layer of insoluble **aluminium oxide (Al_2O_3)** forms on its surface when metal is exposed to air. This layer firmly adheres to the metal and serves to protect the underlying aluminum layers from further corrosion.

On the other hand, the insoluble layer of rust, $Fe_2O_3xH_2O$ that forms on the surface of iron is too porous to protect the underlying metal. This layer flakes away and exposes metal for further corrosion.

Aluminium is extensively used in the construction of aircraft, ships, cars, cooking utensil, window frames, soda canes etc. Aluminium has much higher tendency to oxidize than iron. Therefore, we might expect to see aircrafts, ships, cars, cooking utensils, soda canes transformed into piles of corroded aluminium.

Rusting of iron:

Rusting is a chemical process that, can take place in metals exposed to the atmosphere.

Iron kept in moist air for a long time, gets rusted. The rusting is due to the coating of hydrated ferric oxide on the surface of the metal. Impurities present in the metal facilitate rusting. The presence of CO_2 in air also helps rusting. Along with hydrated oxides rust also consists of carbonate.

Condition for rusting:

Most familiar example of corrosion is the formation of rust on iron. Oxygen and water are necessary for iron to rust. A region of metal surface that has relatively less moisture, acts as anode. Will Fe oxidize in this region?

$$Fe\,(s) \longrightarrow Fe^{+2}(aq) + 2e^-$$

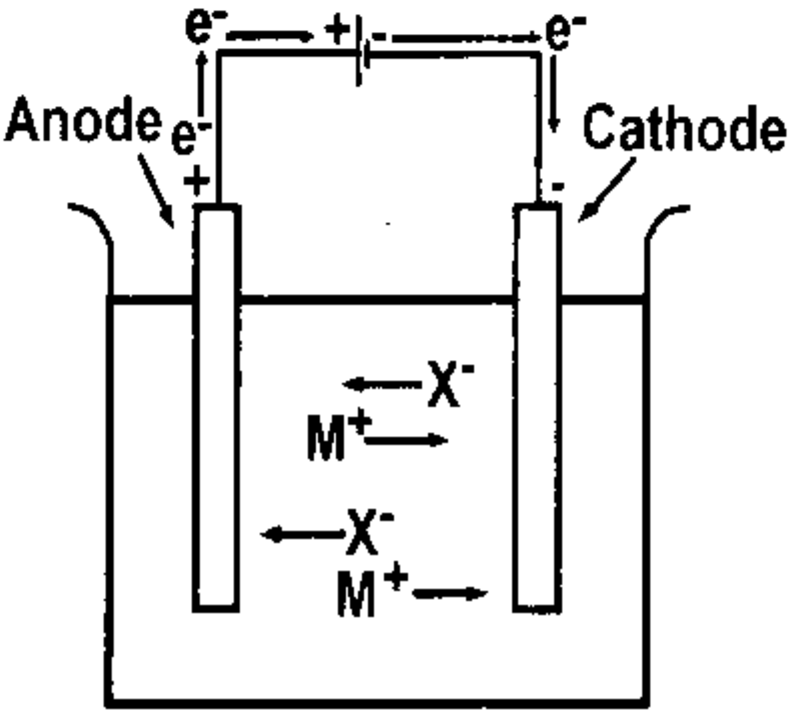
Another region on the surface of metal that has relatively more moisture acts as cathode. The electrons released in the oxidation process reduce atmospheric oxygen to hydroxyl ions.

$$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$$

Q.6 **Electrolysis has a major role in electrochemical industries.**
- **(a)** **Sketch an electrolytic cell, label the anode and cathode and indicate the direction of electron transfer.**
- **(b)** **Describe the nature of electrochemical process.**
- **(c)** **Distinguish between electrolytic and voltaic cell.**

Ans: a. **Electrolytic cells:**

Sketch of electrolytic cell: Figure shows the sketch of an electrolytic cell.



An electrolytic cell

b. **Nature of electrochemical process:**

Electrochemical processes are oxidation-reduction reactions in which chemical energy released by a spontaneous reaction is converted to electricity or in which electrical energy is used to drive a non-spontaneous reaction. Whether an electrochemical process releases or requires energy, it always involves the transfer of electrons from one substance to another. This means that this process always involves an oxidation-reduction or a redox reaction.

c. **Difference between electrolytic and electrochemical cells.**

Electrolytic cell	Galvanic cell
1. In this cell, the electrical energy is converted into chemical energy.	**1.** In this cell, chemical energy is converted into electrical energy.
2. In this cell, current is used to drive a chemical reaction.	**2.** In this cell, current is produced as a result of chemical reaction.
3. Non-spontaneous oxidation-reduction reactions take place.	**3.** Spontaneous oxidation-reduction reactions take place.
4. Electrolysis takes place in this cell.	**4.** Electric conduction takes place in this cell.
5. Examples: Down's cell, Nelson's cell.	**5. Examples:** Daniel's cell, fuel cell.

IMPORTANT SHORT QUESTION ANSWERS (Reduced Syllabus)

Q1. In electroplating of chromium, which salt is used as an electrolyte? Write the redox reaction taking place during the electroplating of chromium.

Ans: Chromium plating:

Since chromium metal does not adhere strongly to the steel therefore steel is first plated with copper or nickel and then chromium.

Electrolyte used:

For electroplating chromium, chromium metal is used as anode and chromium sulphate, $Cr_2(SO_4)_3$ as an electrolyte. A few drops of dil H_2SO_4 are added in the electrolyte to prevent its hydrolysis. The electrolyte ionizes as follows:

$$Cr_2(SO_4)_3\,(aq) \rightleftharpoons 2Cr^{+3}\,(aq) + 3SO_4^{-2}(aq)$$

Redox reactions:

At anode: $Cr\,(s) \rightleftharpoons Cr^{+3}_{(aq)} + 3e^-$

At cathode: $Cr^{+3}_{(aq)} + 3e^- \rightleftharpoons Cr\,(s)$

Chromium plated steel is used to make automobile parts.

Q2. How can a non spontaneous reactions be carried out in an electrolytic cell. Discuss in detail?

Ans: Electrolytic cells:

An electrochemical cell in which electrical energy is used to drive a chemical reaction is called an electrolytic cell.

Construction of an electrolytic cell: An electrolytic cell consists of

(a) A vessel containing an electrolyte (MX) **(b)** Two inert electrodes **(c)** A battery

Working of an electrolytic cell:

The figure shows that electrons move from anode to cathode in the outer circuit, in the solution the cations move towards cathode and anions towards anode. At anode anions oxidize by loosing electrons. At cathode cations reduce by gaining electrons. This means oxidation occurs at anode and reduction at cathode.

Examples of electrolytic cell: Down's cell, Nelson's cell.

Oxidation reaction at Anode: $\overline{X} \rightarrow X + Ie^-$

Reduction reaction at cathode: $M^+ + Ie^- \rightarrow M$

Q3. Give the general principle of a voltaic or galvanic cell. Explain by taking example of Daniel cell?

Ans: Galvanic cells (daniel cells):

The cell which involves spontaneous redox reaction to generates electricity is called a galvanic or voltaic cell. The name Voltaic is given to this cell because Alessandro Volta discovered first such cell.

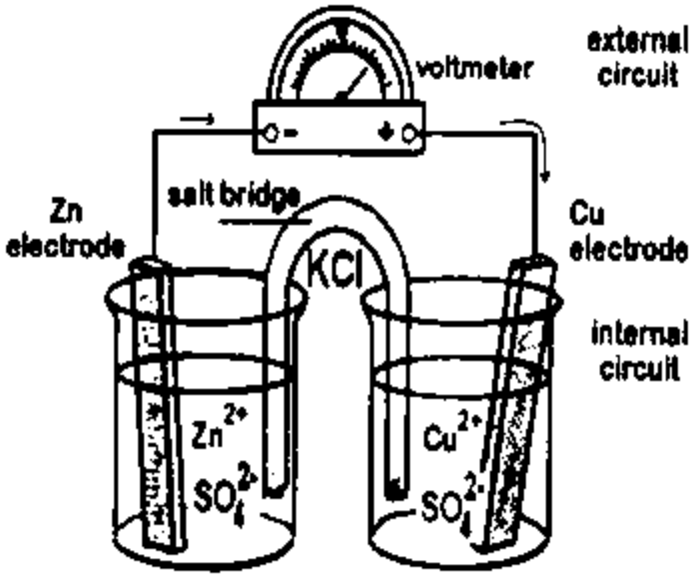
Construction of galvanic cells (daniel cells):

The English chemist Fredrick Daniel constructed first voltaic cell using zinc (Zn) and copper (Cu) electrodes. Therefore this cell is named as Daniel Cell. A galvanic or Daniel cell is shown in figure.

A galvanic cell consists of the following parts:

1. A zinc bar dipped into a 1M Zn SO_4 solution.
2. A copper bar dipped into a 1M Cu SO_4 solution.

4. A voltmeter to measure current.



$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Anode oxidation Cathode reduction

Galvanic cell

Working of Galvanic cells (daniel cells):

Each compartment of cell is called a half cell. Thus a Daniel cell consists of two half cell joined in series. When circuit is complete, electrons begin to flow from Zn rod through the external wires to Cu rod. Thus Zn half cell acts as anode and Cu half cell as cathode. Note that a half cell consists of a metal rod dipped in the solution of its salt.

Reactions in a daniel cell:

In Daniel cell, electrons flow from Zn rod, through the external wire to Cu rod. This is because Zn has more tendency to undergo oxidation than Cu. Zn atoms from the rod go into the solution as Zn^{+2} ions leaving electrons on the rod. These electrons flow in the external circuit. Thus oxidation half reaction occurs at anode compartment. Cu^{+2} ions in copper sulphate solution capture electrons from Cu electrode and are reduced. Reduction half reduction occurs at the cathode compartment. Such oxidation and reduction reactions are called half cell reactions.

At Anode (Oxidation half reaction): $\mathrm{Zn}\,(s) \longrightarrow \mathrm{Zn}^{+2}(aq) + 2e^-$

At Cathode (Reduction half reaction): $\mathrm{Cu}^{+2}(aq) + 2e^- \longrightarrow \mathrm{Cu}\,(s)$

Q4. Find the oxidation state of nitrogen in the following compounds.
(i) NO_2 (ii) N_2O (iii) N_2O_3 (iv) HNO_3

Ans:

i: Oxidation number of N in NO_2: **The sum of oxidation numbers must be zero.** **[O.N of N] + 2[O.N of O] = 0** **[O.N of N] + 2[-2] = 0** **[O.N of N] - 4 = 0** **[O.N of N] = +4** **Thus oxidation state for N in NO_2 is +4.**	**ii: Oxidation number of N in N_2O:** **The sum of oxidation numbers must be zero.** **2[O.N of N] + [O.N of O] = 0** **2[O.N of N] + [-2] = 0** **2[O.N of N] - 2 = 0** **2[O.N of N] = +2** $\frac{2[\text{O.N of N}]}{2} = \frac{+2}{2}$ **[O.N of N] = +1** **Thus oxidation state for N in N_2O is +1.**
iii: Oxidation number of N in N_2O_3:	**iv: Oxidation number of N in HNO_3:**

$2[O.N\ of\ N] - 6 = 0$ $2[O.N\ of\ N] = +6$ $\frac{2[O.N\ of\ N]}{2} = \frac{+6}{2}$ $[O.N\ of\ N] = +3$ Thus oxidation state for N in N_2O_3 is +3.	$[+1] + [O.N\ of\ N] + 3[-2] = 0$ $[+1] + [O.N\ of\ N]\ -6 = 0$ $[O.N\ of\ N] - 5 = 0$ $[O.N\ of\ N] = +5$ Thus oxidation state for N in HNO_3 is +5.

SELF ASSESSMENT EXERCISE 7.2

In the following reactions, identify which element is oxidized and which element is reduced.

i. $Mg + S \longrightarrow MgS$ ii. $4Na + O_2 \longrightarrow 2Na_2O$

iii. $2Al + 3Cl_2 \longrightarrow 2AlCl_3$ iv. $Mg + H_2 \longrightarrow MgH_2$

Solution:

i. $Mg + S \longrightarrow MgS$

$$Mg \longrightarrow Mg^{+2} + 2e^- \text{ (loss of 2 electrons)}$$

Mg atom loses two electrons to form Mg^{+2}, so it is oxidized.

$$S + 2e^- \longrightarrow S^{-2} \quad \text{(gain of 2 electrons)}$$

S atom has gain two electrons to form S^{-2}, so it is reduced.

ii. $4Na + O_2 \longrightarrow 2Na_2O$

$$Na \longrightarrow Na^{+1} + 2e^- \quad \text{(loss of 1 electron)}$$

Na atom loses one electron to form Na^{+1}, so it is oxidized.

$$O + 2e^- \longrightarrow O^{-2} \quad \text{(gain of 2 electrons)}$$

O atom has gain two electrons to form O^{-2}, so it is reduced.

iii. $2Al + 3Cl_2 \longrightarrow 2AlCl_3$

$$Al \longrightarrow Al^{+3} + 3e^- \quad \text{(loss of 3 electrons)}$$

Al atom loses three electrons to form Al^{+3}, so it is oxidized.

$$Cl + 3e^- \longrightarrow 3Cl^{-1} \quad \text{(gain of 3 electrons)}$$

Cl atom has gain three electrons to form $3Cl^{-1}$, so it is reduced.

iv. $Mg + H_2 \longrightarrow MgH_2$

$$Mg \longrightarrow Mg^{+2} + 2e^- \text{ (loss of 2 electrons)}$$

Mg atom loses two electrons to form Mg^{+2}, so it is oxidized.

$$H + 2e^- \longrightarrow 2H^{-1} \quad \text{(gain of 2 electrons)}$$

H atom has gain two electrons to form $2H^{-1}$, so it is reduced.

GUESS PAPER & MODEL PAPER # 8
BASED ON CHAPTER # 8 (Reduced Syllabus)
CHEMICAL REACTIVITY

CHAPTER 8 : CHEMICAL REACTIVITY

Metal, electropositive character, comparison of reactivity of Alkali and alkaline Earth metals, inertness of Noble metals, importance of silver, Gold, Platinum, Non-metals, Electro negativity character, comparison of reactivity of halogen.

Note: Topic related self-assessments, review exercise and think tank questions are included. Side boxes and Society, science and technology are not included.

SECTION-A

Time allowed: 20 Minutes **Marks: 12**

Note: Section-A is compulsory. All parts of this section are to be answered on the question paper itself. It should be completed in the first 20 minutes and handed over to the Centre Superintendent. Deleting/overwriting is not allowed. Do not use lead pencil.

Q.1 Encircle the correct option i.e. A / B / C / D. All parts carry equal marks.

(i) Which of the following oxides is the most basic oxide?
A. MgO B. BaO C. CaO D. SrO

(ii) Which of the following oxides is amphoteric oxide?
A. Na_2O B. Al_2O_3 C. MgO D. SO_2

(iii) Which of the following elements is most metallic?
A. Al B. Na C. Mg D. Li

(iv) Valence shell electronic configuration of alkaline earth metals is:
A. ns^1 B. ns^2np^1 C. ns^2 D. ns^2np^6

(v) Which of the following metals is not noble metal?
A. Au B. Cu C. Hg D. Ag

(vi) Which of the following elements is the least non-metallic?
A. Si B. S C. P D. Cl

(vii) Oxidizing power of ________ is the highest.
A. I_2 B. Cl_2 C. Br_2 D. F_2

(viii) Chlorine cannot oxidize ________ ion.
A. Iodide B. Flouride C. Bromide D. Astatine

(ix) Na is more reactive than Li, but less reactive than ____________.
A. Mg B. Al C. Rb D. All of these

(x) Which of the following is the strongest acid?
A. HI B. HCl C. HF D. HBr

(xi) Iodine (I_2) is _________ solid.
A. Pale yellow B. greenish yellow C. reddish brown D. bluish black

(xii) In group IIA, the basic character of oxides increases in the following order.
A. $MgO < BeO < CaO < SrO < BaO$ B. $BeO < MgO < CaO < SrO < BaO$
C. $BeO < MgO < CaO < BaO < SrO$ D. $BaO < CaO < MgO < SrO < BeO$

Time allowed: 2:40 hours **Total Marks: 53**

Note: Answer any six parts from Section 'B' and attempt any five parts from Section-C. Attempt any two questions from Section 'D' on the separately provided answer book. Use supplementary answer sheet i.e. Sheet–B if required. Write your answers neatly and legibly.

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. $(6 \times 3 = 18)$

i. In a group the reactivity of metals with oxygen increases? Give example.
ii. Which element is more metallic Mg or Al? Explain.
iii. What is the importance of Mg?
iv. Arrange the following in order of increasing acidic strength. HF, HI, HBr, HCl
v. List the important properties of metals?
vi. Why it is advisable, not to pick sodium metal with fingers.
vii. Arrange the following elements in order of increasing non-metallic character.
Si, Al, P.
viii. Explain the electropositive character of the metals?

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. $(5 \times 3 = 15)$

i. What are the characteristics of electronegative elements and how do their characters vary in groups and periods?
ii. Highlight the important characteristics of some important alkaline earth metals?
iii. Arrange the following oxides in order of decreasing basic character.
BeO, CaO, MgO, SrO
iv. Rank the each set of elements in order of increasing metallic character.
(a) Al, Na, Mg (b) Na, Li, K
v. Give some important applications of platinum.
vi. What do you mean by inertness of noble metals?
vii. What is aqua regia?

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. $(2 \times 10 = 20)$

Q.4 a. Which of the following displacement reaction will not occur? Give reason.
(a) $Cl_2 + 2NaF_2 \longrightarrow 2NaCl + F_2$ (b) $Br_2 + 2KI \longrightarrow 2KBr + I_2$
(c) $I_2 + 2KBr \longrightarrow 2KI + Br_2$
b. Highlight the Position of alkali metals and alkaline earth metals in the period table?

Q.5 a. Describe the Ionization energies of alkali and alkaline earth metals?
b. Highlight the important characteristics of some important alkali metals?

Q.6 a. What do you know about halogens?
b. How do you compare the reactivity of halogens?

SOLUTION OF GUESS PAPER & MODEL PAPER # 8
(Reduced Syllabus)

SECTION-A (MCQs)

i. B	ii. B	iii. A	iv. C	v. C	vi. A

SECTION – B (Marks 18)

Q.2 Attempt any SIX parts from the following. All parts carry equal marks. (6 × 3 = 18)

i. In a group the reactivity of metals with oxygen increases? Give example.

Ans: In a group the reactivity of metals with oxygen increases For instance, alkali metals on exposure to air show an increase in affinity for oxygen as we move down the group. Lithium forms normal oxide.

$$4Li(s) + O_2(g) \longrightarrow 2Li_2O(s)$$

Sodium forms per oxide. $2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$

K, Rb and Cs form super oxide $K(s) + O_2(g) \longrightarrow KO_2(s)$

In periods as we move from left to right, reactivity of metals with oxygen decreases.

Examples:

For instance, in the third period sodium readily reacts with oxygen to form sodium peroxide, at room temperature. Mg reacts with oxygen only on ignition. Al reacts with oxygen on heating. It is superficially oxidized at room temperature to form Al_2O_3 layer which serves as a protective covering that prevents further reaction.

ii. Which element is more metallic Mg or Al? Explain.

Ans: Mg and Al are the members of period 3. Apply the trends of decreasing electropositive character across a period. Check the periodic table to see which elements is more metallic.
Mg is more metalic than Al.

iii. What is the importance of Mg?

Ans: Importance of magnesium:

Magnesium has relatively low density and it is used in making light weight alloy with aluminum that are used for making frames of automobiles , air craft and space ship cameras etc.

Since magnesium burns brilliantly, it is used in photo flash guns. $Mg(OH)_2$ called milk of magnesia is commonly used as antacid.

iv. Arrange the following in order of increasing acidic strength.
HF, HI, HBr, HCl

Ans: The hydrogen halides dissolve in water to form hydrohalic acid e.g., hydrochloric acid HCl, hydrofluoric acid HF etc. Except HF, other hydrohalic are strong acids. The acidic strength increases in the following order:

$$HF < HCl < HBr < HI$$

v. List the important properties of metals?

Ans: Metals:

Metals like aluminum, iron and copper can be transformed into a variety of shapes by melting, molding, casting, hammering and rolling. Metals occupy left and lower three-quarters (3/4) of the periodic table. They are typically shiny solids, have moderate to high melting points and are good conductor of heat and electricity. They tend to lose electrons in chemical reactions.

vi. Why it is advisable, not to pick sodium metal with fingers.

Ans: Reaction of sodium with water:

Sodium reacts violently with water (moisture of fingers), often catching fire and exploding.
This reaction produces sodium hydroxide, a strong and caustic base.

vii. Arrange the following elements in order of increasing non-metallic character.
Si, Al, P.

Ans: Since Si, Al and P lies in period 3 and in period from left to right non metallic character increases therefore:

$$P < Si < Al$$

viii. Explain the electropositive character of the metals?

Ans: Electropositive character (Electropositivity):

Where M stands for any metal. Metals have large atomic size and low ionization energies.

Variation of electropositive character in groups:

As ionization energy decreases down the group, the electropositivity increases. Thus sodium is more electropositive than lithium. Similarly, magnesium is more electropositive than beryllium.

Variation of electropositive character in periods:

The tendency to lose electron decreases as we move from left to right in a period. For instance in the second period Li and Be are metals, B is metalloid, whereas C, N, O, F and Ne are non-metals.

Be is less metallic than Li. These trends in electropositivity are reflected in chemical reactivity of metals.

SECTION – C (Marks 15)

Q.3 Attempt any FIVE parts from the following. All parts carry equal marks. $(5 \times 3 = 15)$

i. What are the characteristics of electronegative elements and how do their characters vary in groups and periods?

Ans: Electronegative or non-metallic character:

Electronegative character is the reverse of electropositive character. It is the tendency to gain electrons to form anions.

OR

The tendency of an element to accept an electron to form an anion is called its non-metallic or electronegative character.

$$X(g) + 1e^- \longrightarrow X^-(g)$$

Characteristics of electronegative elements:

The elements having **high electron affinity** or high electronegativity have higher tendency to gain electrons and form **negative ion.** The elements in the **upper right hand portion** of the periodic table are electronegative or non-metallic in nature. For example, fluorine, oxygen and phosphorus are non-metallic in nature.

Greater the tendency to form an anion greater is the non-metallic character of the element.

Variation of electronegative character in groups:

In groups, electronegative characters decreases from top to bottom due to increases in the atomic size.

Variation of electronegative character in periods:

In periods, electronegative characters increases from left to right due to decreases in the size of atoms.

Note: Halogens are the most electronegative elements in their respective periods. Non-metal oxides are generally acidic since they yield acids in water. Acidic strength of these oxides increases from left to right in a period.

ii. Highlight the important characteristics of some important alkaline earth metals?

Ans: Some important alkaline earth metals:

Magnesium and calcium are the most common alkaline earth metals. Metallic bonding in these metals involve two electrons. Therefore these metals are relatively soft but are much harder than alkali metals. They are relatively reactive but less reactive than alkali metals.

Magnesium reacts with steam but calcium reacts with cold water.

$Mg(s) + H_2O(g) \longrightarrow MgO + H_2(g)$

$Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2 + H_2(g)$

Magnesium reacts with oxygen at high temperature whereas calcium reacts with oxygen at room temperature.

$$2Mg(s) + O_2(g) \xrightarrow{heat} 2MgO(s)$$

$$2Ca(s) + O_2(g) \xrightarrow{heat} 2CaO(s) + O_2(g)$$

Hence Ca is more reactive than Mg.

Uses of magnesium:

Magnesium has relatively low density and it is used in making light weight alloy with aluminum, that are

Uses of calcium:

Calcium is an important component of teeth and bones Calcium ions are found in all living cells. They are involved in clotting of blood. A proper Ca^{+2} **and** K^{+} ions balance is required for normal heart function.

iii. Arrange the following oxides in order of decreasing basic character.
BeO, CaO, MgO, SrO

Ans: Basic character of metal oxides of main group elements increases down the group. For example, in group IIA, the basic character of oxides increases in the following order.

$$SrO > CaO > MgO > BeO$$

iv. Rank the each set of elements in order of increasing metallic character.
(a) Al, Na, Mg (b) Na, Li, K

Ans: a. Al, Na and Mg lies in period 3. As in period from left to right electropositive character decreases therefore decreasing metallic character is given by. **Na < Mg < Al**

b. Na, Li and K lies in group IA. As in group from top to bottom electropositive character increases therefore decreasing metallic character is given by. **Li < Na < K**

v. Give some important applications of platinum.

Ans: Important applications of platinum:

Platinum is widely used as catalyst for many types of industrial processes. For example, 100% pure sulphuric acid is prepared by the contact process. In this process platinum is used as catalyst Platinum is used as electrode as a part of hydrogen electrode and in fuel cells. Automobile exhaust is a major source of air pollution. Therefore, most new cars are equipped with catalytic converters. These converters contain platinum. These converters catalyze the complete combustion of CO and hydrocarbons. A platinum compound called cis-platin is useful as an anti cancer agent.

vi. What do you mean by inertness of noble metals?

Ans: Inertness of noble metals:

The chemistry of metals is characterized by their ability to lose electrons to form cations. Some metals such as copper, silver, gold and platinum are relatively difficult to oxidize. Therefore, these metals are often called noble metals Gold and platinum exist mostly as free elements in nature. Copper and silver exist in both free and combined states.

All active metals react with HCl but noble metals do not react with HCl. Copper and silver react with strong oxidizing agents such as conc. HNO_3 and $HClO_4$.

Aqua regia:

Gold and platinum react only with aqua regia. Aqua regia is a mixture of 3 parts by volume of conc. HCl and one part by volume of conc. HNO_3.

vii. What is aqua regia?

Ans: Aqua regia (Royal water):

Gold and platinum react only with aqua regia. Aqua regia is a mixture of 3 parts by volume of conc. HCl and one part by volume of conc. HNO_3.

SECTION – D (Marks 20)

Note: Attempt any TWO questions. All questions carry equal marks. **($2 \times 10 = 20$)**

Q.4 a. Which of the following displacement reaction will not occur? Give reason.

(a) $Cl_2 + 2NaF_2 \longrightarrow 2NaCl + F_2$ (b) $Br_2 + 2KI \longrightarrow 2KBr + I_2$

(c) $I_2 + 2KBr \longrightarrow 2KI + Br_2$

Solution:

(a) $Cl_2 + 2NaF_2 \longrightarrow 2NaCl + F_2$

Displacement reaction will not occur because the reactivity of halogens decrease down the group Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$. Therefore Cl_2 cannot oxidized F_2.

(b) $Br_2 + 2KI \longrightarrow 2KBr + I_2$

Displacement reaction will occur because the reactivity of halogens decrease down the group. Thus,

Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$ Therefore I_2 cannot oxidized Br_2.

b. Highlight the Position of alkali metals and alkaline earth metals in the period table?

Ans: Position of alkali metals and alkaline earth metals in the period table:

The group IA elements except hydrogen are known as alkali metals. These include lithium, sodium, potassium rubidium, cesium and francium. These elements have general electron configuration ns^1 in their valence shell. The Group IIA elements are known as alkaline earth metals. These include beryllium, magnesium calcium, strontium, barium and radium. They have general electron configuration ns^2 in their valence shell.

Q.5 a. Describe the Ionization energies of alkali and alkaline earth metals?

Ans: Ionization energies of alkali metals:

Because alkali metals have low ionization energies, therefore have a great tendency to lose the single valence electron to form cations. They are so reactive that they are never found in the free state. The alkali metals show increased reactivity down the group. This is because of decrease in ionization in energies down the group. This means Na is more reactive than Li, but less reactive than K. They are all so reactive that they are kept under a liquid such as kerosene oil.

Table	
Li 520	Be 899
Na 495	Mg 738
K 419	Ca 590
Rb 403	Sr 549
Cs 376	Ba 503

Ionization energies of alkaline earth metals:

Alkaline earth metals also have low ionization energies, so they also have great tendency to lose both the valence electrons to form dispositive cations. They are less reactive than alkali metals. These metals also show increased reactivity down the group. Thus Mg is more reactive than Be but less reactive than Ca. Like alkali metals, they are also kept under a liquid such as kerosene oil to prevent contact with moist air.

b. Highlight the important characteristics of some important alkali metals?

Ans: Some important alkali metals:

The most important alkali metal is sodium. It is hard and strong. So when the term metal comes in our mind, we think it to be as hard as iron. But sodium is a soft metal can be cut with a knife. It melts at 97.8°C and boils at **881.4°C**. The reason is that in iron, metal atoms are tightly packed due to small size. They are held through strong metallic bonds due to many electrons in the valence shell Whereas sodium consists of bigger atoms. Because of large size and only one electron, sodium atoms have weak metallic bonds. Thus sodium is a soft metal. Sodium is a silvery white soft metal. It is an extremely reactive metal

Reaction of sodium with water:

Sodium reacts with water violently to form sodium hydroxide and liberates hydrogen gas.

$$2Na(s) + H_2O(l) \longrightarrow 2NaOH\,(aq) + H_2(g)$$

This reaction is highly exothermic and proceeds with light explosion. For this reason, it should not be picked with fingers.

Reaction of sodium with oxygen:

In a limited supply of oxygen sodium burns to form sodium oxide (Na_2O). But in excess of oxygen it forms pale yellow solid sodium per oxide (**Na_2O_2**). $2Na(s) + O_2(g) \longrightarrow 2Na_2O_2(g)$

Uses of sodium:

Q.6 a. What do you know about halogens?

Ans: Halogens:

The elements in group VII-A are called halogens. The name halogen is derived from the Greek words "halous" meaning salt and "gen" meaning former. Halogens include fluorine, chlorine, bromine, iodine and astatine (astatine is a radioactive element. Little is known about its properties) All halogens are reactive non-metals. They all exist as diatomic molecular substances.

Fluorine (F_2) ⟶ Pale yellow gas
Chlorine (Cl_2) ⟶ Greenish yellow gas
Bromine (Br_2) ⟶ Reddish-brown liquid
Iodine (I_2) ⟶ Bluish-black solid

All the halogens react with metals to form salts called halides. In these reactions metals are oxidized. So, halogens act as oxidizing agents. Fluorine is the most reactive element known. Chlorine is less reactive, but combines vigorously with many metals. For instance sodium metal burns in chlorine gas to form sodium chloride

$$2Na\,(s) + Cl_2\,(g) \longrightarrow 2NaCl\,(s)$$

Bromine and iodine react with metals less vigorously.

b. How do you compare the reactivity of halogens?

Ans: Comparison of reactivity of halogens:

The comparison of **reactivity of halogen** shows that it decreases down the group.

All the halogens are the most reactive elements as a family. Fluorine is the most reactive due to its high electronegativity, small size, extremely high oxidizing power and low dissociation energy of F-F bond.

The reactivity decreases with atomic number, e.g., in the displacement reactions we find that fluorine displaces all the remaining three halogens from their salts Chlorine displaces only bromine and iodine, bromine displaces only iodine while iodine displaces none.

Bromine and iodine react with metals less vigorously The reactivity of halogens decrease down the group. Thus, the order of decreasing power oxidizing is:- **$F_2 > Cl_2 > Br_2 > I_2$**

The halogens react with hydrogen to form hydrogen halides. $H_2(g) + X_2(g) \longrightarrow 2HX(g)$

Where **X= F, Cl, Br, I**

Reaction of fluorine with hydrogen is explosive. The hydrogen halides dissolve in water to form hydrohalic acid e.g., hydrochloric acid HCl. hydrofluoric acid HF etc. Except HF, other hydrohalic are strong acids. The acidic strength increases in the following order: HF < HCl < HBr < HI

Oxidizing power of F_2 is the highest and that of I_2 is lowest. Due to the relative strength as oxidizing agent, it is possible for a free halogen to oxidize the ion of halogen next to it in the group. This means F_2 can oxidize all the halide ions to free halogen.

For example,

$$F_2(g) + 2KCl(aq) \longrightarrow 2KF(aq) + Cl_2(g)$$

$$F_2(g) + 2KBr(aq) \longrightarrow 2KF(aq) + Br_2(l)$$

$$F_2(g) + 2NaI(aq) \longrightarrow 2NaF(aq) + I_2(s)$$

Similarly Cl_2 can oxidize Br^- and I^- ions. But I_2 cannot oxidize any halide ion.

IMPORTANT SHORT QUESTION ANSWERS
(Reduced Syllabus)

Q1. Compare and contrast alkali and alkaline earth metals.

Ans: Alkali metals similarities Alkaline earth metals:

iv. In flame both group elements give different colors .
v. Both group elements soluble in liquid ammonia and give deep blue solution
vi. Both are present in s-block .
vii. Both are react with H and forms respective hydrides.
viii. In both grope differentiating electron enters in s sub shell.
ix. Both group elements have electro positive character

Difference between alkaline earth metals and alkali metals:

	Properties	Alkali metals	Alkaline earth metals
1.	Physical properties	Soft, low melting points paramagnetic.	. Comparatively harder. High melting points. Diamagnetic.
2.	Valency	Monovalent	Bivalent
3.	Electropositive nature	More electropositive	Less electropositive
4.	Hydroxides	Strong bases, highly soluble and stable towards heat.	Weak bases, less soluble and decompose on heating.
5.	Bicarbonates	These are known in solid state.	These are not known in free state. Exist only in solution.
6.	Carbonates	Soluble in water. Do not decompose on heating ($LiCO_3$ is an exception)	Insoluble in water. Decompose on heating.
7.	Action of carbon	Do not directly combine with carbon	Directly combine with carbon and form carbides
8.	Solubility of salts	Sulphates, phosphates, fluorides, chromates, oxides etc are soluble in water.	Sulphates, phosphates fluorides, chromates, oxalates etc are insoluble in water
9.	Reducing power	Stronger as ionization potential values are low and oxidation potential values are high	Weaker as ionization potential values are high and oxidation potential values are low.
10.	Electronic configuration	One electron is present in the valency shell. The configuration is ns^1 (monovalent) more electropositive	Two electrons are present in the valency shall. The configuration is ns^2 (bivalent)

Q2. Choose three metals that would not be suitable for coinage. Elaborate your choice.

Ans: Group One elements are known as the alkali metals. They are all silvery-white in colour, react vigorously with water, very soft, and react rapidly with air, and are thus totally unsuitable for coins. The majority of these metals are either too reactive, too expensive or just too difficult to work to be used in coins and medals.

For Example: Lithium, Sodium, Potassium, Rubidium and Caesium.

SELF ASSESSMENT EXERCISE 8.1

1. **In each of the following pairs identify, which element is more metallic?**
   (a) Be, B **(b)** Si, Al **(c)** K, Li
2. **Using the periodic table, rank each set of elements in order decreasing metallic character**
   (a) Na, Li, K **(b)** Al, Na, Mg
3. **Which is basic Li_2O or Na_2O?**

Solution:

1. **In each of the following pairs identify, which element is more metallic?**
   (a) Be, B **(b)** Si, Al **(c)** K, Li

b. Apply the trends of increasing electropositive character down a group and decreasing electropositive character across a period.

Si and Al lies in period 3. As in period from left to right electropositive character decreases therefore Al is more metallic element (Al proceeding Si in period 3).

c. Apply the trends of increasing electropositive character down a group and decreasing electropositive character across a period.

K and Li lies in group IA. As in group from top to bottom electropositive character increases therefore K is more metallic element (Li proceeding K in group IA).

2. Using the periodic table, rank each set of elements in order decreasing metallic character

(a) Na, Li, K **(b)** Al, Na, Mg

Ans: a. Na, Li and K lies in group IA. As in group from top to bottom electropositive character increases therefore decreasing metallic character is given by. **K > Na > Li**

b. Al, Na and Mg lies in period 3. As in period from left to right electropositive character decreases therefore decreasing metallic character is given by. **Na > Mg > Al**

3. Which is basic Li_2O or Na_2O?

Ans: Li and Na lies in group IA. As in group from top to bottom basic character of metal oxide increases therefore Na_2O is more basic than Li_2O.

SELF ASSESSMENT EXERCISE 8.2

Identify the position of Mg and Ca in the periodic table.

Ans: Magnesium is group IIA element is known as alkaline earth metal. Magnesium has general electron configuration ns^2 in its valence shell.

Calcium is group IIA element is known as alkaline earth metal. calcium has general electron configuration ns^2 in its valence shell.

SELF ASSESSMENT EXERCISE 8.3

1. In each of the following pairs, identify which element is less non-metallic.

(a) B or C **(b)** C or Si **(c)** Cl or Br.

2. Using periodic table rank each set of elements in order of increasing non-metallic character.

(a) N, F, O **(b)** Cl, Br, I **(c)** Si, S, P

Solution: 1. In each of the following pairs, identify which element is less non-metallic.

(a) B or C **(b)** C or Si **(c)** Cl or Br

Ans: a. Since B and C lies in period 2 and in period from left to right non metallic character increases therefore B is less non metallic than C.

b. Since C and Si lies in group IVA and in group from top to bottom non metallic character decreases therefore Si is less non metallic than C.

c. Since Cl and Br lies in group VIIA and in group from top to bottom non metallic character decreases therefore Br is less non metallic than Cl.

2. Using periodic table rank each set of elements in order of increasing non-metallic character.

(a) N, F, O **(b)** Cl, Br, I **(c)** Si, S, P

Ans: a. Since N, F and O lies in period 2 and in period from left to right non metallic character increases therefore: N < O < F

b. Since Cl, Br and I lies in group VIIA and in group from top to bottom non metallic character decreases therefore: Cl < Br < I

SELF ASSESSMENT EXERCISE 8.4

Which of the following displacement reactions will occur?

1. $Cl_2(g) + 2NaF(aq) \longrightarrow 2NaCl(aq) + F_2(g)$
2. $Br_2(g) + 2KI(aq) \longrightarrow 2KBr(aq) + I_2(g)$
3. $I_2(g) + 2KBr(aq) \longrightarrow 2KI(aq) + Br_2(l)$
4. $Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(l)$
5. $Cl_2(g) + 2NaI(aq) \longrightarrow 2NaCl(aq) + I_2(s)$

Solution:

1. $Cl_2(g) + 2NaF(aq) \longrightarrow 2NaCl(aq) + F_2(g)$

Displacement reaction will not occur because the reactivity of halogens decrease down the group. Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$
Therefore Cl_2 cannot oxidized F_2.

2. $Br_2(g) + 2KI(aq) \longrightarrow 2KBr(aq) + I_2(g)$

Displacement reaction will occur because the reactivity of halogens decrease down the group. Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$
Therefore Br_2 can oxidized I_2.

3. $I_2(g) + 2KBr(aq) \longrightarrow 2KI(aq) + Br_2(l)$

Displacement reaction will not occur because the reactivity of halogens decrease down the group. Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$
Therefore I_2 cannot oxidized Br_2.

4. $Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(l)$

Displacement reaction will occur because the reactivity of halogens decrease down the group. Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$
Therefore Cl_2 can oxidized Br_2.

5. $Cl_2(g) + 2NaI(aq) \longrightarrow 2NaCl(aq) + I_2(s)$

Displacement reaction will occur because the reactivity of halogens decrease down the group. Thus, the order of decreasing power oxidizing is: $F_2 > Cl_2 > Br_2 > I_2$
Therefore Cl_2 can oxidized I_2.